

KEYS TO ENVIRONMENTAL CHEMISTRY

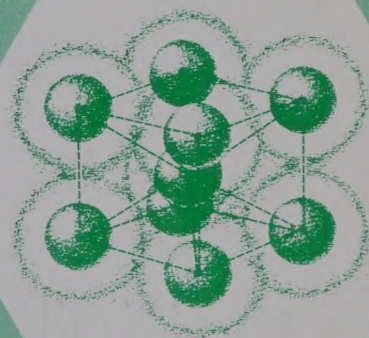
University Of Alberta



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A Module in the KEYS TO CHEMISTRY Program



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RATIONALE FOR MODULES

The core program of KEYS TO ENVIRONMENTAL CHEMISTRY is contained in the textbook. However, it was a part of this program that supplemented the core program developed to do the following: (1) to provide for the needs of the interested and capable student; (2) to provide remedial work for students who might be deficient in basic skills; and (3) to offer a variety of options for students in the fields of environmental chemistry, practical applications of chemistry in daily living, and interesting by-ways which might lead to hobbies and/or avocations that are chemistry-related. By selecting the modules that fit his or her own goals and interests, the student can plan a truly individualized chemistry course.

THE KEYS TO ENVIRONMENTAL CHEMISTRY MODULES are patterned after the chapters of the core textbook. However, because of the character of this module it does not contain laboratory experiments and self-test items. It is anticipated that the student will use the KEYS TO ENVIRONMENTAL CHEMISTRY module to investigate the condition of the local environment. A report of findings would result which can be used to evaluate the student's effort. This module fits into the third category of modules. It may be studied after completion of Chapter 10 of the core textbook.

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KEYS TO ENVIRONMENTAL CHEMISTRY

LUCY T. PRYDE • JAY A. YOUNG • ELAINE W. LEDBETTER

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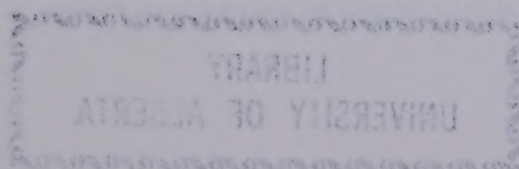


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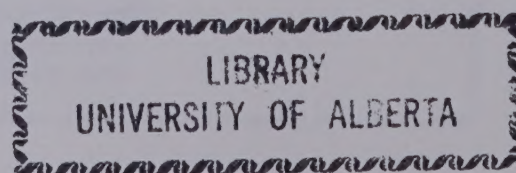
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INTRODUCTION

This module provides you with an opportunity to study experimentally two resources that are very important in determining environmental quality—air (the atmosphere) and water (the hydrosphere). Even though these two resources are vital to your well-being, both are commonly used as dumping grounds for various environmental contaminants, or pollutants. (*Pollutants* are materials or energy forms—i.e. thermal or nuclear radiation present in the environment in sufficient quantities to be harmful to living things.)

The atmosphere and the hydrosphere are both very large, constantly changing systems. Their size and movement led many people to believe that air and water are unlimited in their capacity to absorb pollutants. Slowly people have come to realize that this is not the case. Instead, the interrelationships among all the different parts of the environment have become more apparent at every turn. For example, solid particles emitted, or given out, by an industry or an automobile can be carried long distances by air and water currents. As a result, particles of lead have been found in the snow of the High Sierras, far out in the oceans, and even in Antarctica. There is really no complete disposal of the particles, only recycling of them in the environment.

Chemistry has a major role to play in helping people understand and solve environmental problems. Many of the problems become clear once the basic chemical principles involved are understood. Chemistry also can provide the means to measure and evaluate the effects of chemicals in air and water. These measurements and evaluations are important pieces of information to have when making plans for improving the quality of the environment. Although many other aspects of human society are involved, chemistry provides one important tool for understanding the environment and for improving environmental quality.

How you will use this module is up to you. You may plan to spend all of your available time in one of the two major areas, thereby allowing yourself time for carrying out some of the suggestions for further study given with each experiment. On the other hand, you may plan to do a few experiments from each part to gain an overview of air and water quality. However you use this module, you will find it a useful key in unlocking some of the many complex problems related to environmental quality.

OBJECTIVES FOR KEYS TO THE ENVIRONMENT: AIR AND WATER QUALITY

By the time you have completed your study of both parts of this module, you will be expected to demonstrate acceptable performance on the following learning objectives:

1. State the expected composition of pure dry air and pure water.
2. List several common air and water contaminants.
3. Discuss the sources and significance of each contaminant you listed.
4. Measure the concentrations of several substances found in air and water.
5. Explain several methods for purifying air and water.
6. Identify air and water-quality problems that may occur in your area and discuss possible solutions.

Note: Specific objectives for investigating air quality and water quality are listed at the beginning of Parts One and Two, respectively, along with a *Suggested Order of Study* for that part. Either part may be done first. You may want to review the introductory material on pages 1–4 of Laboratory Keys to Chemistry. Such a review will refresh your memory on how to work in the laboratory effectively and safely.

PART ONE INVESTIGATING AIR QUALITY

65,000 km

OBJECTIVES

By the time you have completed Part One of this module, you will be expected to demonstrate acceptable performance in the following learning objectives:

1. State the normal mixture of gases in the lower atmosphere.
2. List several processes that cause variations in the local air environment.
3. Name several chemical substances that are considered to be air pollutants.
4. Measure the particulate dustfall in your area and relate it to air quality.
5. Discuss how the blackness of smoke is related to the quantity of particulate matter in the smoke.
6. Convert gas concentration expressed in percent into parts per million and vice versa.
7. Measure the concentration of carbon monoxide in air and discuss its relationship to air quality.
8. Detect the presence of sulfur dioxide in the atmosphere and relate it to air quality.
9. Demonstrate the effect that air pollution can have on materials.
10. Discuss the possible waste products from burning plastics.
11. Summarize the test you have performed as indicators of air quality.
12. Discuss the process of filtration as a method for purifying air.
13. Explain how an electrostatic precipitator is useful in removing particulate matter from an air stream.
14. Demonstrate and explain the process of wet scrubbing for the removal of gaseous contaminants from air.
15. Summarize the methods of air purification you have tested.

SUGGESTED ORDER OF STUDY

1. Study Environmental Section 1-1.
2. Environmental Experiments 1-1 through 1-6 deal with air pollutants and how they affect materials. Read through these experiments and select the ones you wish to carry out. Perform the experiments you selected in numerical order. Hand in a report for each experiment.
3. Environmental Experiments 1-7 through 1-9 deal with the purification of air. Read through them and decide which ones you wish to carry out. Again perform the ones you chose in numerical order. Hand in a report for each experiment.
4. Check the items listed in the section entitled *Suggestions for Further Study* provided at the end of each experiment. Use the items to plan and carry out some creative work.
5. Explore the outside references listed in the *Bibliography* at the end of Part One.
6. Check with your teacher for evaluation of your learning experience with Part One of this module.

ENVIRONMENTAL SECTION 1-1 • INVESTIGATING THE ATMOSPHERE

The earth is surrounded by a mixture of gases called the *atmosphere*. Liquid and solid particles are suspended in the atmosphere. Only a small portion of the total atmosphere directly concerns you (see Fig. 1). This portion is the lower atmosphere, also called the *troposphere*. The troposphere contains about 75% of the

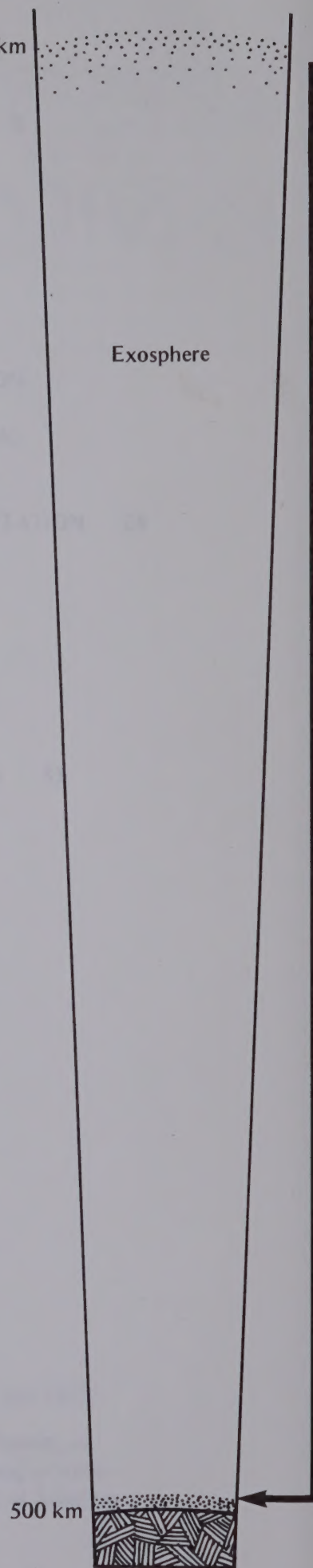
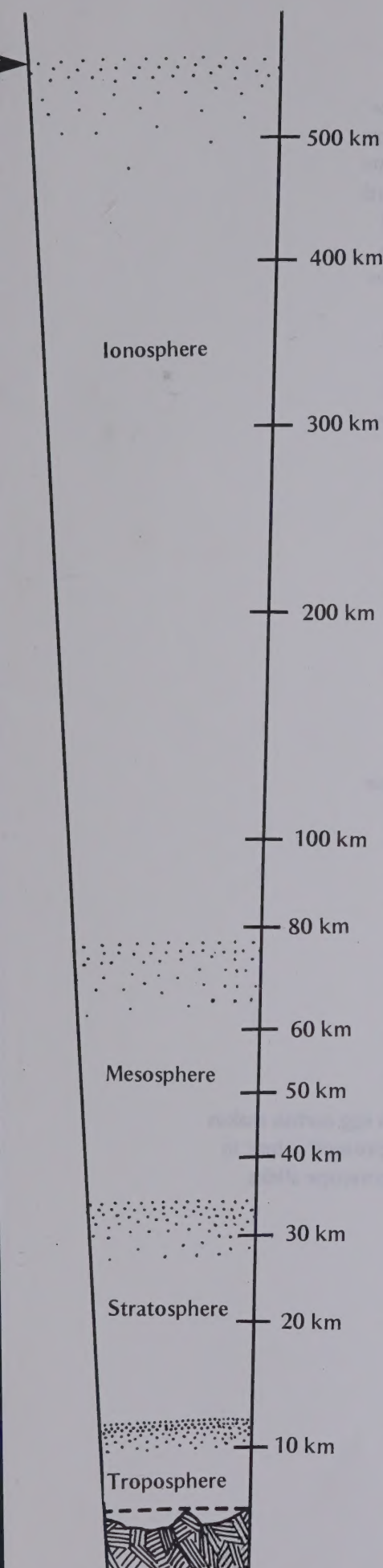


FIGURE 1



total amount of air and nearly all of the water vapor in the entire atmosphere. It is the region of weather, including winds, rain, snow, and clouds. In general, the temperature drops with increasing height in the troposphere, except for localized temperature inversions.

The atmosphere can be considered to be composed of a mixture of gases. They can be grouped as pure dry air, water vapor, and pollutants. By volume, pure dry air consists of a mixture of about 78% nitrogen, about 21% oxygen, about 0.9% argon, about 0.03% carbon dioxide, and even smaller percentages of neon, helium, krypton, xenon, and several other gases. The relative amount of each gas in pure dry air varies little. In contrast, the amount of water vapor in the atmosphere varies widely from time to time and from place to place.

To this basic mixture of pure dry air and water vapor, natural processes and human activities add a wide variety of pollutants. Some of them are small particles, such as dust, and others are gases. For example, in the United States alone, over 100 million tons per year of carbon monoxide is emitted. Some of it comes from the burning of fuel for transportation in cars, trucks, buses and locomotives and for power in electric generating plants. Additional carbon monoxide gas comes from industrial processes, disposal of solid waste, and miscellaneous sources. Other gaseous emissions that create air pollution are sulfur dioxide, hydrocarbons, and nitrogen oxides. In addition, many particles, larger than gas molecules, are produced. They add to the general pollution.

What happens to pollutants once they are in the lower atmosphere? Small amounts of them are rapidly dispersed by moving air. As a result, their local effects are diluted. Some of the larger particles settle out quickly, while others remain suspended in the atmosphere for years. Besides reducing visibility, an accumulation of suspended particles in the atmosphere may significantly affect the global climate by scattering incoming radiation. When activated by the energy of sunlight, certain pollutants react with other components of the atmosphere to form photochemical smog. The air around you is in a constant state of change. In fact, the atmosphere really serves as a giant reaction vessel for a vast series of physical and chemical changes.

The quality of air in the rather thin slice of life-supporting atmosphere should be of concern to everyone. People depend on the purity of the air for life itself and can ill afford to allow its composition to be significantly changed. Many people have already discovered that air pollution affects both health and the economy, as well as being unpleasant to see and smell. In this part of the module you will be investigating ways to test the quality of air and various methods for purifying air.

ENVIRONMENTAL EXPERIMENT 1-1 • MEASUREMENT OF DUSTFALL

Purpose

To measure dustfall.

Materials

- | | |
|--|--------------------------------------|
| 1 knife or pair of scissors | 1 balance, sensitive to a milligram |
| 1 egg carton or other carrying case | 1 small box with lid |
| 1 glass-marking pencil | aluminum foil, cardboard, tacks, and |
| 4 microscope slides | twine or string (optional) |
| 1 15-cm ruler | 1 cc acetone |
| two-sided silicone tape or silicone grease | 1 magnifying glass or microscope |

Discussion

Tiny solid and liquid particles suspended in the atmosphere are referred to as *particulate air pollution*. The size of the particles helps to determine their fate in the atmosphere. Particles which are relatively large (greater than 10^{-5} m) usually settle out due to gravity. In addition, rain is helpful in removing particles larger than 10^{-6} m. Particles smaller than 10^{-6} m, however, are light enough to be carried by air currents and stay suspended in the atmosphere for long periods of time. As an example, the fine dust from the 1883 eruption of Krakatoa (an Indonesian volcanic island) remained in the atmosphere for over a year, causing beautifully colored sunrises and sunsets all over the world. In this experiment you will be collecting particles heavy enough to settle out of the atmosphere during a 30-day period.

There are many sources of solid particles in the air other than volcanoes. Some natural ones are dust from dry soils, unburned materials from forest fires, and pollen. Human activities also add solid particles to the atmosphere. Examples are burning of fuels, construction and demolition projects, industrial grinding and pulverizing operations, and mining processes.

Inhalation of finely divided particles is known to cause medical problems, since these small particles can hamper the proper functioning of the lungs. Respiratory diseases such as emphysema and chronic bronchitis may result. In addition to the particles themselves being hazardous, they may carry with them adsorbed gases and liquids that can cause further damage once they are in the lungs.

Procedure

1. Cut the egg carton so it can carry 3 microscope slides without touching each other and without anything touching the tops of the slides (Fig.2).

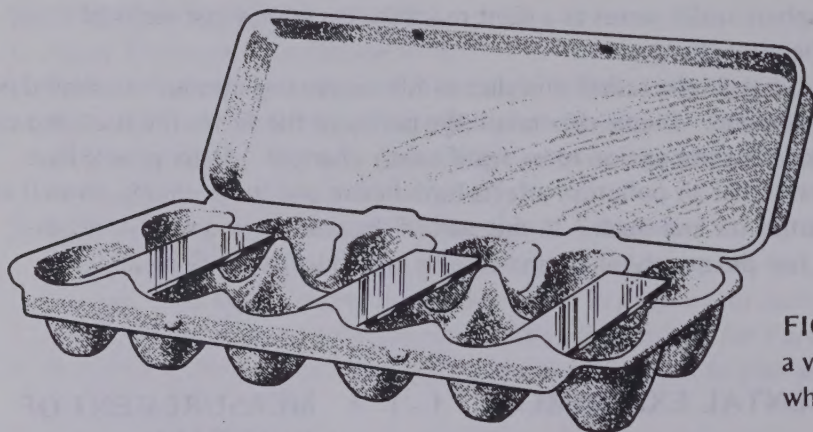


FIGURE 2 An egg carton makes a very convenient protective box in which to carry microscope slides.

2. With a glass-marking pencil, add an identifying code to each of the microscope slides. Measure and record the length and width of each slide to the nearest tenth of a centimeter. Calculate the surface area in square centimeters.
3. Wash and dry each slide; do not disturb the identifying code. Cover the top of each slide with a layer of two-sided silicone tape or silicone grease applied in a thin, even layer. From now on, avoid touching the tops of the slides.
4. Measure and record the mass of each slide to the nearest milligram.
5. Place one of the slides in a labeled, covered container and leave it in a safe location in the laboratory where it will not get excessively warm and will not be tampered with accidentally. This slide will serve as a control and will not be exposed to air currents for the 30-day period.
6. Take the other 3 microscope slides home in the carrying box. Place each slide, sticky side up, in a different location. Choose locations where the slides

will be exposed to air currents and yet protected from people, pets, heat, moisture, and direct sunlight.

Among the possible locations indoors are these: the top of a bookshelf, a window ledge, and the top of a refrigerator. You may, if you wish, suspend a slide from a ceiling (see Fig. 3). If you give the slides some protection,

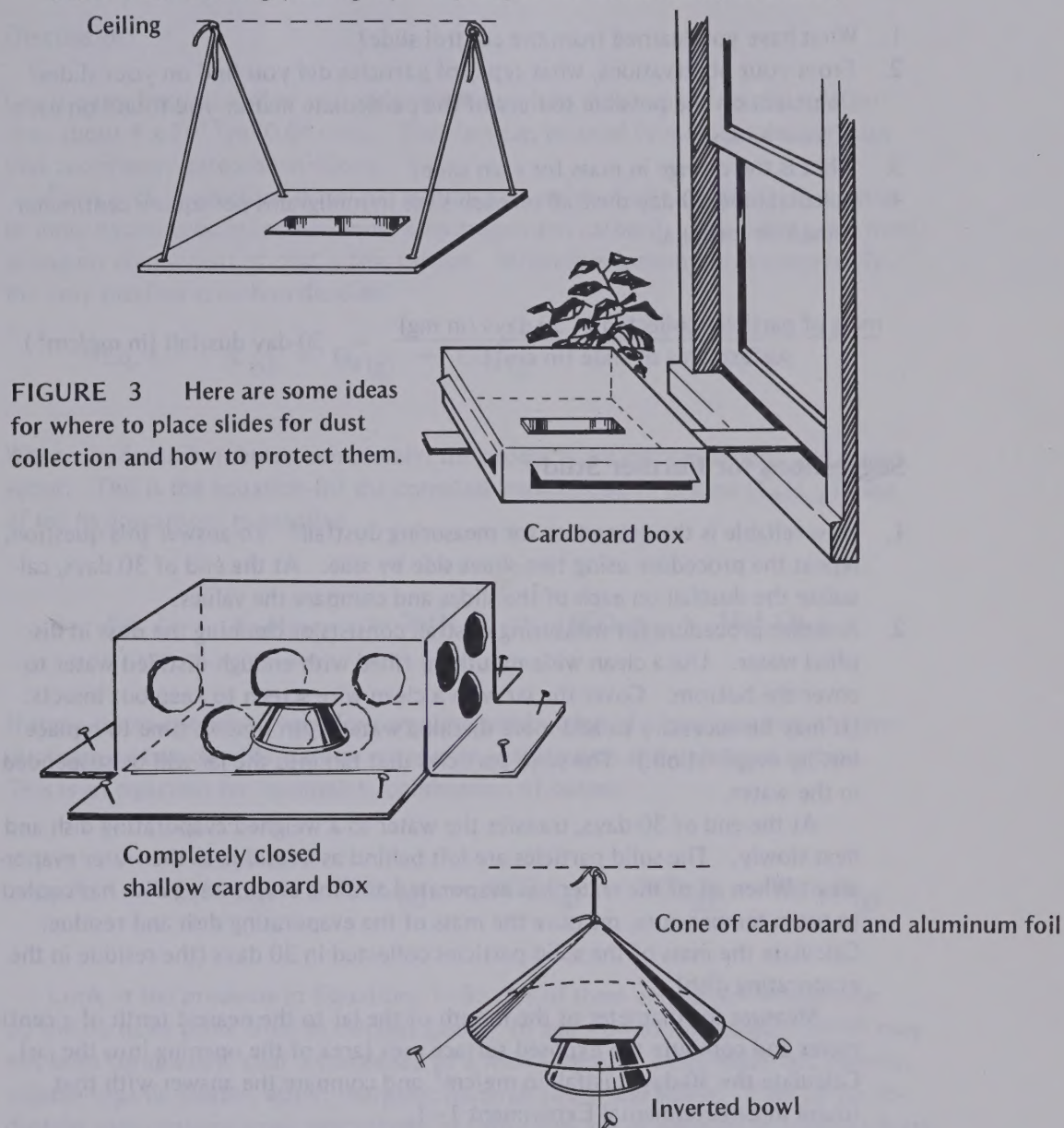


FIGURE 3 Here are some ideas for where to place slides for dust collection and how to protect them.

you may choose locations outdoors. In Fig. 3 are some suggestions for protecting your slides that others have used successfully. Undoubtedly, you can think of other arrangements, too. Remember that air must be able to circulate freely over each slide. If needed, you can anchor a slide on the bottom with two-sided cellophane tape. Before reweighing the slide, you can clean off the gummy material with acetone.

Record the code on each slide along with the location of the slide and the method, if any, used to protect it.

7. After a period of 30 days, carefully pick up each of the 3 slides and put them in the carrying case. Take the carrying case with the slides in it back to the laboratory. Clean off the bottom of each slide, using acetone if necessary. Measure and record the mass of each slide to the nearest milligram. Also measure and record the mass of your control slide.
8. Examine each slide through a magnifying glass or microscope. Make as many visual observations as possible of the color, size, and types of particulate

matter you have collected. Record your observations of each slide along with the code on the slide.

Questions

1. What have you learned from the control slide?
2. From your observations, what types of particles did you find on your slides? Comment on the possible sources of the particulate matter you found on each of your slides.
3. What is the change in mass for each slide?
4. Calculate the 30-day dustfall on each slide in milligrams per square centimeter using this formula.

$$\frac{\text{mass of particles collected in 30 days (in mg)}}{\text{surface area of slide (in cm}^2\text{)}} = \text{30-day dustfall (in mg/cm}^2\text{)}$$

Suggestions for Further Study

1. How reliable is the procedure for measuring dustfall? To answer this question, repeat the procedure using two slides side by side. At the end of 30 days, calculate the dustfall on each of the slides and compare the values.
2. Another procedure for measuring dustfall consists of catching the dust in distilled water. Use a clean wide-mouth jar filled with enough distilled water to cover the bottom. Cover the jar with a clean wire screen to keep out insects. (It may be necessary to add more distilled water from time to time to replace loss by evaporation.) The solid particles that fall into the jar will be suspended in the water.

At the end of 30-days, transfer the water to a weighed evaporating dish and heat slowly. The solid particles are left behind as a residue as the water evaporates. When all of the water has evaporated and the evaporating dish has cooled to room temperature, measure the mass of the evaporating dish and residue. Calculate the mass of the solid particles collected in 30 days (the residue in the evaporating dish).

Measure the diameter of the mouth of the jar to the nearest tenth of a centimeter and compute the exposed surface area (area of the opening into the jar). Calculate the 30-day dustfall in mg/cm² and compare the answer with that found in Environmental Experiment 1-1.

Why must the water used be distilled? What advantages or disadvantages are there in this method compared with the first?

ENVIRONMENTAL EXPERIMENT 1-2 • DETERMINATION OF SHADE OF SMOKE

Purpose

To determine shade of smoke as an indication of particulates in exhaust gases.

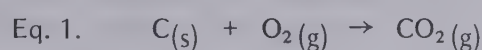
Materials

The chart in Figure 4

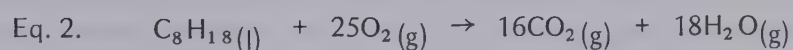
Discussion

Gas streams become visible as clouds or smoke when they contain particles larger than about 4×10^{-5} m (0.04 mm). This fact can be used to monitor the particles that accompany gaseous emissions.

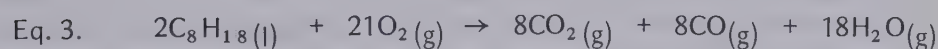
Fuels such as gasoline, diesel oil, natural gas, and fuel oil consist mainly of one or more hydrocarbons (compounds of hydrogen and carbon). In contrast, the most abundant component of coal is free carbon. When free carbon burns completely, the only product is carbon dioxide.



When a hydrocarbon burns completely, the products are carbon dioxide and water vapor. This is the equation for the complete combustion of octane (C_8H_{18}), one of the hydrocarbons in gasoline.



If there is not enough oxygen for complete combustion of a hydrocarbon, the products are carbon dioxide, carbon monoxide, water vapor and unburned carbon. This is an equation for incomplete combustion of octane.



Look at the products in Equations 1-3. All of them are gases. Where then do the particles come from in exhaust gases? In addition to free carbon, which may not burn completely, coal is composed of a wide variety of other materials, namely, volatile organic matter, water, inorganic minerals (ash), and sulfur. Fuel oil for industrial uses contains small percentages of sulfur and ash in addition to hydrocarbons. Fuel oil for domestic use and diesel oil also contain sulfur and ash but in smaller percentages than industrial fuel oil. In contrast, natural gas consists of only hydrocarbons, and so does gasoline unless small amounts of additives have been added. These additional materials found in fuels are also involved in the process of combustion. As a result, sulfur dioxide and fly ash (noncombustible mineral particles that range in size from 10^{-3} mm to 10^{-1} mm) may be emitted with the stream of exhaust gases. Furthermore, very few processes of combustion actually burn a fuel completely. Therefore, the products of combustion will often include particles of unburned fuel and particles of partially burned fuel.

Unburned carbon particles are thought to be primarily responsible for the gray or black color of exhaust gases. The mixture of unburned carbon particles and exhaust gases is known as *smoke*. If one assumes that the shade of the smoke emitted is directly related to the concentration of unburned carbon particles in the smoke, he or she can use the shade of smoke as an indication of the completeness of combustion or the efficiency of the combustion process.



FIGURE 4 Smoke density chart adapted from Ringelmann Smoke Chart (courtesy of U.S. Bureau of Mines)

Procedure

1. Examine the smoke density chart (Figure 4). It consists of a set of standard grids composed of black lines on a white background. Hold the chart at arm's length. Now the grids appear as graduated shades of gray.
2. Locate several places where smoke is being discharged. Check factories, houses, power plants, and construction sites as possible sources of smoke. Determine which shade of gray, or numbered grid, on the smoke density chart most closely matches the shade of smoke at each source you located. For each source, record its type, its location, the distance you estimate you are from it at the time you determine the shade of smoke, and the number of the grid that most nearly matches the shade of smoke. Also record any other pertinent observations you make of each source.
3. Estimate the density of smoke at each source you chose. The density of smoke represented by grid No. 1 is 20%; by grid No. 2, 40%; by grid No. 3, 60%; and by grid No. 4, 80%. Dense, black, smoke can be considered to have a density of 100%.

Questions

1. What process might be responsible for the smoke you observed at each of the places you selected?
2. What additional information would you need if you wished to use the shade of smoke for regulating air pollutions?
3. If a smokestack or exhaust pipe emits only billows of "white smoke", what conclusion can you draw about the combustion of hydrocarbons taking place? About the amount and type of pollutants in the "white smoke?"
4. If no visible gases are coming from an exhaust pipe or smokestack, what conclusion can you draw about the combustion of hydrocarbons taking place? About the amount and type of pollutants in the exhaust gases?

Suggestions for Further Study

1. Visit, call, or write your regional Air Pollution Control District to learn how the density of smoke test is used in your area to help control air pollution.
2. If the engine of an automobile is properly tuned, no black smoke will come

out of the tailpipe when the engine is running. What kind of combustion is occurring in an engine that is producing black smoke? What needs to be done to the engine to eliminate the black smoke?

- Air pollution from particulate matter could be reduced greatly if all processes of combustion were discontinued. What alternatives are there to the combustion of carbon and hydrocarbons for producing electrical power? For driving an automobile? How practical is each of the alternatives at the present time? Does each alternative also have one or more undesirable environmental effects?

ENVIRONMENTAL EXPERIMENT 1-3 • MEASUREMENT OF CARBON MONOXIDE

Purpose

To measure the concentration of carbon monoxide in the atmosphere.

Materials

- | | |
|-----------------------------------|-----------------------|
| 1 carbon monoxide test kit | 1 Celsius thermometer |
| 3 carbon monoxide indicator tubes | |

Discussion

In the last experiment you learned that the process of incomplete combustion can add solid particles to the atmosphere. You also learned that the simplest oxide of carbon, carbon monoxide (CO), is formed when the carbon or hydrocarbons in fuels burn incompletely. Carbon monoxide is normally present in the atmosphere at a concentration of only 0.00001% by volume. This is equivalent to 0.1 part per million (ppm).

Part per million is a common unit used in reporting the concentration of a gas in air. The unit is particularly suitable for expressing the concentration of any gas present in air in relatively small amounts.

It is easy to convert percent by volume to parts per million, for percentages are parts per hundred. For example, neon is present normally in the atmosphere at a concentration of 0.0018% by volume, or 0.0018 parts per hundred, which can be written 0.0018/100. To convert parts per hundred to parts per million, set up a proportion in which the letter x equals the number of parts of neon in a million parts of air and determine the value of x .

$$\frac{0.0018}{100} = \frac{x}{1,000,000}$$

$$x = \frac{0.0018 \times 1,000,000}{100}$$

$$x = 18$$

Thus, 0.0018 parts per hundred equals 18 ppm. This is equivalent to multiplying parts per hundred by 10,000 (10^4). To multiply by 10,000, move the decimal four places to the right.

Many processes add carbon monoxide to the atmosphere. In urban areas, various forms of transportation account for over 90% by weight of all carbon mon-

oxide emissions. Nationwide, approximately 78% of all carbon monoxide emissions are from transportation. The rest of the carbon monoxide in the atmosphere comes from industry, power plants, solid-waste disposal, and miscellaneous processes. On the global scale, most (80%) of the carbon monoxide in the atmosphere comes from oxidation of methane (CH_4) emitted by decaying organic matter.

Eventually, carbon monoxide in the atmosphere is oxidized to carbon dioxide with the help of hydroxyl free radicals (OH^\cdot) and soil bacteria. (A *free radical* is a reactive molecular fragment which has an unpaired electron.) Were this natural mechanism for removing carbon monoxide from air not at work, carbon monoxide would accumulate rapidly in the environment.

Why should carbon monoxide emissions be a problem? Because carbon monoxide is colorless and odorless, exposure to it can take place without a person's even being aware of the presence of the gas. Unlike most air pollutants, carbon monoxide does not irritate the lungs. Instead, it mimics the affinity of oxygen for the hemoglobin of the red blood cells, attaching itself to hemoglobin two hundred times more strongly than oxygen. Therefore, oxygen cannot be properly distributed to body tissues when carbon monoxide is present in the lungs. Even at rather low concentrations, prolonged exposure to carbon monoxide can result in headache, nausea, and dizziness. These and other reactions to carbon monoxide are given in Figure 5. Since hemoglobin will regain its capacity to carry oxygen if the person moves into an atmosphere free of carbon monoxide, the effects of the gas are usually reversible.

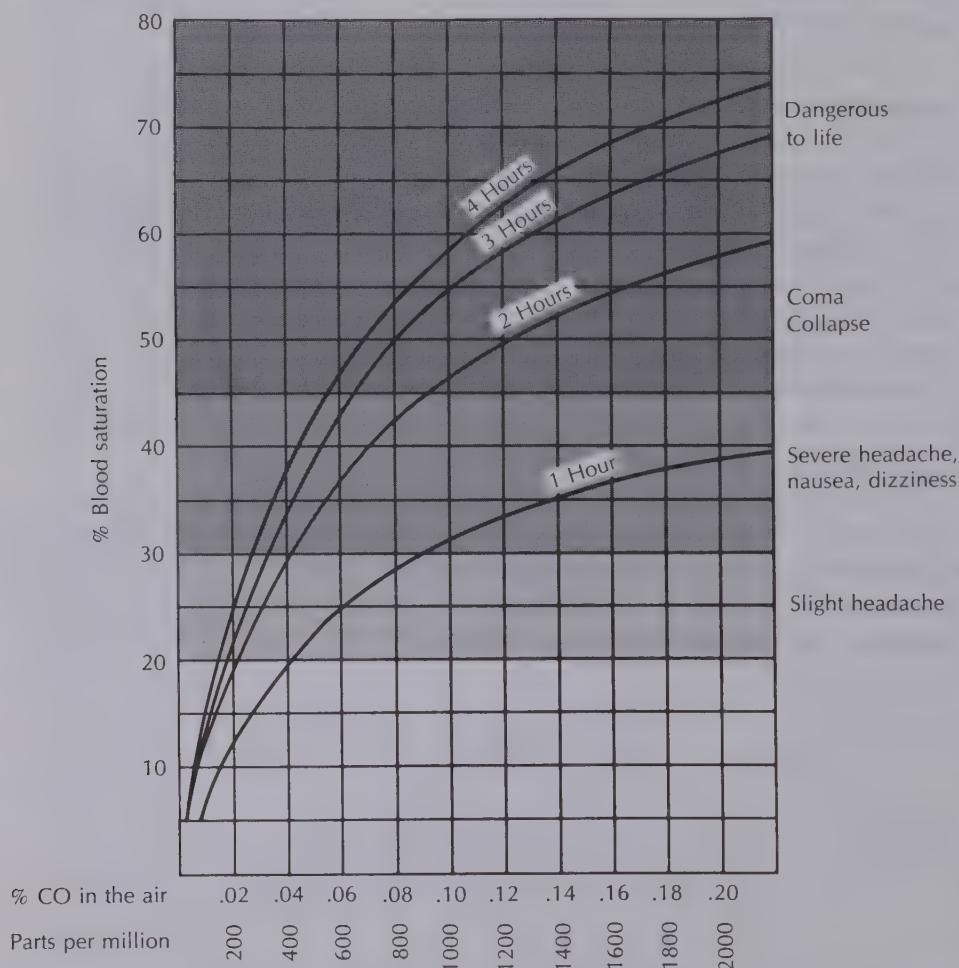


FIGURE 5 Effects of Carbon Monoxide on Humans
(courtesy of Mine Safety Company)

Procedure

1. Examine the carbon monoxide test kit. Note that it consists of an indicator tube and a pump to draw air through the tube. Read and follow all directions given with the kit. In particular, remember to eject all air from the pump before putting the indicator tube in place. Otherwise you are simply testing the air inside the pump rather than the gas you intended to test.
2. Select three different places where you can test air for the presence of carbon monoxide. Some suggested places are these: inside a running automobile standing outdoors, at the end of the tailpipe of a running automobile parked outdoors (SAFETY NOTE: Never touch the tailpipe of a car whose engine has been running--the engine generates tremendous heat which can cause the tailpipe to become very hot), near an industrial furnace, at a bus terminal, in a tunnel, at the entrance of a manhole, and above a charcoal or wood fire. Very likely, you can think of other places, too.
3. Test a sample of air at each of the places you selected. When a sample contains a detectable concentration of carbon monoxide, that is, a concentration greater than 0.001%, or 10 ppm, the center of the indicator tube will turn from yellow to green. The color formed is due to a two-step chemical process in which the carbon monoxide reacts with ammonium molybdate in the presence of a catalyst containing palladium to form a complex molybdate ion. The material formed is actually blue in color, but in the presence of the yellow starting materials, it appears green.
4. If you detect the presence of carbon monoxide, determine its concentration by matching the intensity of the green color to the color key in the tester. This will give you a quantitative measure of the carbon monoxide present. At concentrations above 1,000 ppm, a difference in the green color is no longer produced.

Questions

1. What was the concentration of carbon monoxide in each of the places where you tested? In each case give the possible source or sources of the carbon monoxide you detected.
2. Write a chemical equation for the incomplete combustion of pure carbon. *Hint:* carbon reacts with oxygen gas to yield carbon monoxide.
3. Currently in the U.S. the maximum allowable concentration of carbon monoxide in the atmosphere is 9 ppm for 8 hours of exposure, with a maximum concentration of 35 ppm for 1 hour of exposure. How does each of your measurements compare with this standard? Would there be adverse effects from remaining in any of the air you tested?

Suggestions for Further Study

1. Carbon monoxide is one of the products when cigarettes or other products containing tobacco are burned. Considering the low burning temperature, it is not surprising that incomplete combustion takes place, producing carbon monoxide. One way to test the concentration of carbon monoxide in cigarette smoke is to set up a "smoking machine" such as the one shown in Figure 6.

Turn on the vacuum pump or faucet aspirator and light the cigarette. When the flask is full of cigarette smoke turn off the vacuum pump or faucet aspirator. Test the smoke in the flask for the presence of carbon monoxide. If you detect carbon monoxide, determine its concentration. How does the concentration of carbon monoxide, if detected in the smoke, compare with the federal standard

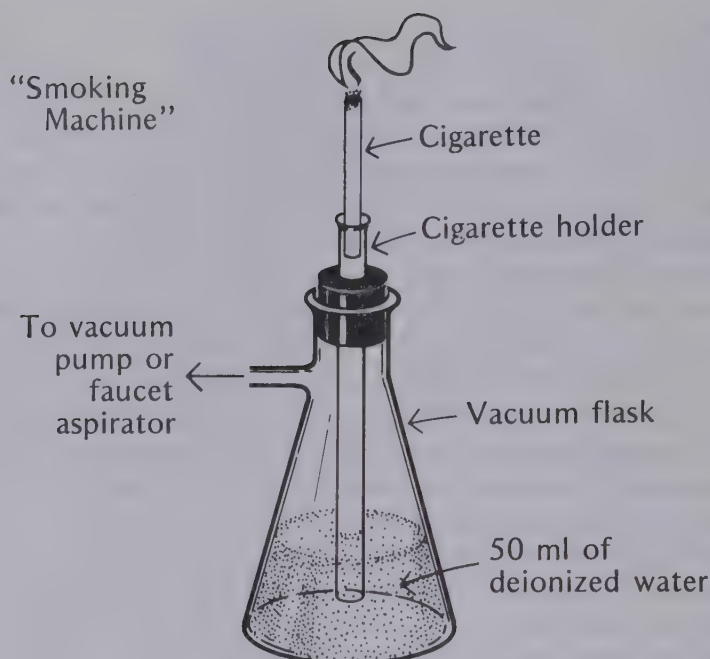


FIGURE 6 A smoking machine.

for concentration of carbon monoxide in the atmosphere? Would there be adverse effects from remaining in this concentration of carbon monoxide?

2. What is the maximum allowable concentration of carbon monoxide in the atmosphere for your state? How does this concentration compare with the federal regulation?
3. What are the allowable carbon monoxide emissions from automobiles? What methods are used to test for compliance with this standard for automotive emissions?

ENVIRONMENTAL EXPERIMENT 1-4 • DETECTION OF SULFUR DIOXIDE

Purpose

To detect the presence of sulfur dioxide in the atmosphere.

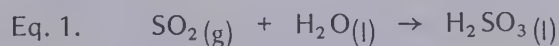
Materials

- | | |
|--|--|
| 1 side-arm test tube or filter flask | 1 faucet aspirator or vacuum pump |
| 1 one-hole rubber stopper to fit side-arm test tube | aluminum foil |
| 1 right-angle glass bend with one end tapered to form a nozzle | 2 test tubes |
| 1 10-cc graduated cylinder | 1 cork or solid rubber stopper to fit test tube |
| 10 cc sodium tetrachloromercurate solution | 10 cc standard solution containing the equivalent of 0.14 ppm sulfur dioxide |
| 2 pieces rubber tubing, thick walled | 1 cc pararosaniline solution |
| 1 flow meter (0 to 5 liters/min) | 1 cc formaldehyde solution |

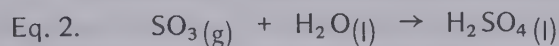
Discussion

Sulfur dioxide (SO_2) is an important gas to monitor in the air because it can have far-reaching effects on vegetation, human beings and materials, generally. The normal concentration of sulfur dioxide in air is only 0.0002 ppm, but the combustion of sulfur-containing coal, industrial fuel oil, diesel oil, and gasoline adds sulfur dioxide to the atmosphere. In addition, industrial operations such as smelters can emit large amounts of sulfur dioxide. Solid-waste disposal and to a small extent the burning of fuel oil for domestic use add sulfur dioxide to the atmosphere as well. In a chemical laboratory, the concentration of sulfur dioxide can become high when students perform experiments involving the burning of sulfur or the generation of sulfur dioxide as a product of some chemical oxidation-reduction reactions.

Most of the effects of sulfur dioxide can be linked with its ability to combine readily with water in the air to form a weak acid, namely, sulfurous acid (H_2SO_3).

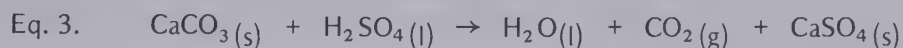


Sulfur dioxide is converted to sulfur trioxide (SO_3) rather slowly; however, the reaction is catalyzed by particulates such as those in fly ash. The sulfur trioxide, once formed, combines with water in the air to form a strong acid—sulfuric acid (H_2SO_4).



Both sulfurous and sulfuric acids are suspended in the atmosphere as tiny droplets and may wash down during the first rain. The use of sulfur-containing fuels may lead, therefore, to some rather unpleasant effects when they are added to the atmosphere by the combustion of these fuels.

Because carbon dioxide and water in the air react to form the weak acid, carbonic acid (H_2CO_3), normally rain might be expected to have a pH of 5.7. However, pH values of 2.1 for rainwater have been recorded during storms in the United States. In urban and surrounding areas of the Northeast where most pH measurements of rain have been taken, the average pH readings of rain are about 4. This acid rain can corrode steel, destroy fabrics and leather, and dissolve limestone and marble, which are both forms of calcium carbonate.



This reaction has caused extreme damage to statues and public buildings, including the Lincoln Memorial in Washington, D.C. One official compared the Lincoln Memorial in the rain to a giant Alka-Seltzer tablet.

Sulfur dioxide has severe effects on vegetation. A splotchy, bleached pattern on its leaves is usually evidence of damage to plants by sulfur dioxide. Apple and pear trees are particularly susceptible to sulfur dioxide and show damage even after a low dose, such as 0.5 ppm of sulfur dioxide for 6 hours. Human beings, too, are susceptible to sulfur dioxide. The established relationship between concentrations of sulfur dioxide and respiratory symptoms has caused the World Health Organization to rate sulfur dioxide as the most dangerous air pollutant. When both sulfur dioxide and particulates are present, the damage is particularly severe. The federal standard for concentration of sulfur oxides in the atmosphere is 0.14 ppm for 24 hours and 0.03 ppm as an annual average.

Procedure

1. Set up the apparatus as shown in Figure 7.

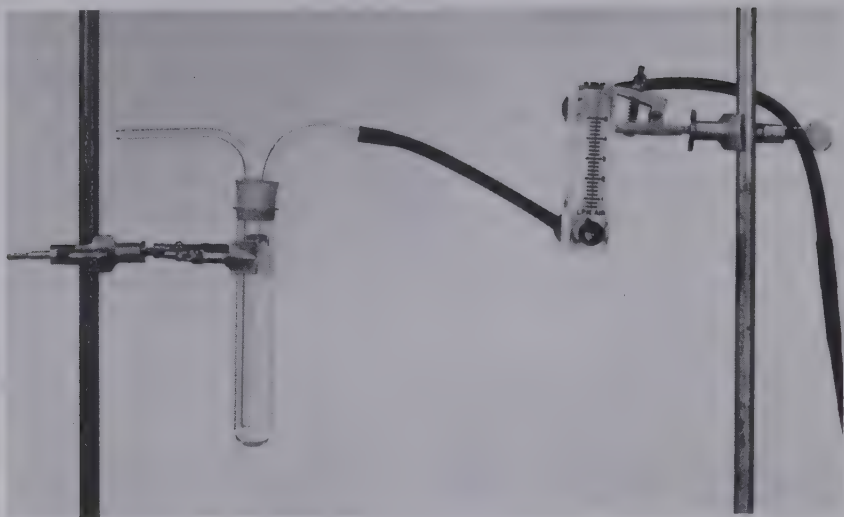


FIGURE 7 Apparatus for metering gas flow.

2. Wrap the side-arm test tube in aluminum foil to protect the solution in it from strong light. Draw 10 liters of air through the solution in the side-arm test tube. (*Hint:* If you can adjust the flow rate to 1 liter/min, you will need to operate the apparatus for 10 min.) If there is sulfur dioxide in the air, it will react with the tetrachloromercurate ion to yield a stable product.
3. Transfer the material from the side-arm test tube into one of the regular test tubes. Put 10 cc of the standard sulfur dioxide solution in the other test tube. Add 1 cc of pararosaniline solution and 1 cc of formaldehyde solution to the material in each test tube. Be careful in handling these solutions; do not inhale any of their fumes. Be sure the materials in the test tubes are mixed. Stopper each test tube and wrap them in aluminum foil. Allow both test tubes to stand for 15 min or until no further increase in color is noted. The formation of a reddish-purple color indicates the presence of sulfur dioxide. The depth of the color is proportional to the concentration of sulfur dioxide present in the sampled air and the standard solution.

Questions

1. Does the amount of sulfur dioxide in the air sampled fall above or below 0.14 ppm?
2. If you detected sulfur dioxide in the air you sampled, what is the most likely source of it?
3. Write an equation for the combustion of sulfur to produce sulfur dioxide.
Write an equation for the conversion of sulfur dioxide to sulfur trioxide.

Suggestions for Further Study

1. To test samples of air in locations where you do not have access to a vacuum pump or water aspirator, you can draw air through a sodium tetrachloromercurate solution by setting up an arrangement like the one in Figure 8. You can also use this arrangement for the main experiment if you do not have access to a flow meter. Use a large glass or plastic container such as a carboy used to ship a large amount of liquid. It must have a capacity larger than 10 liters. Add exactly 1 liter of water to the container and mark the upper level of the water on the side of the container. Repeat the procedure until you can no

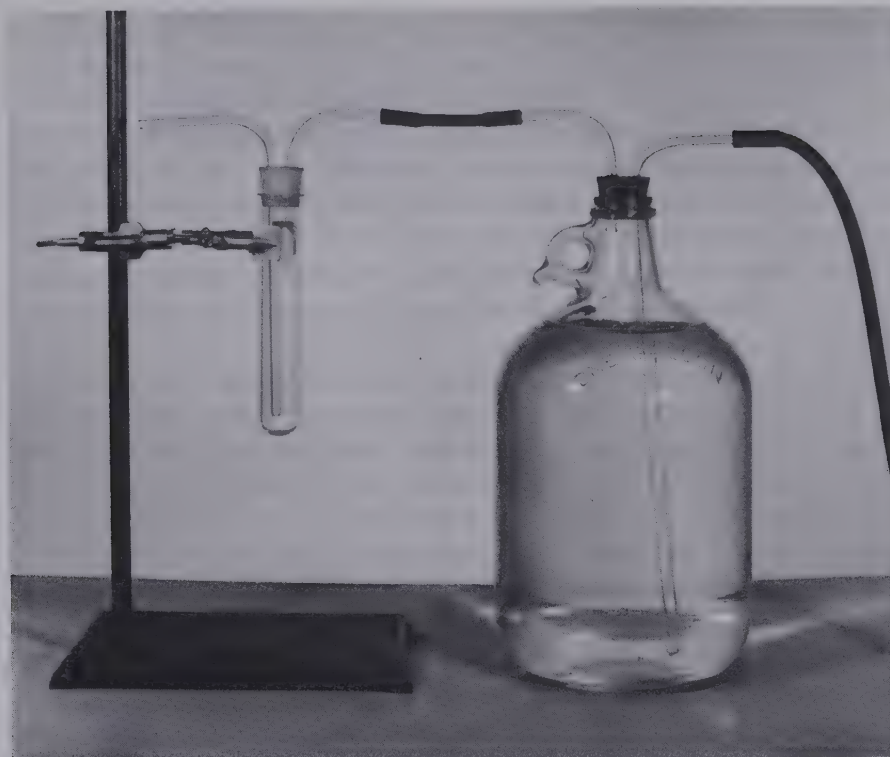


FIGURE 8 Another setup for metering gas flow.

longer add a whole liter of water to the container. Number the lines on the side of the container from 0 to 10, starting with the top line.

Open the pinch clamp and siphon off 10 liters of water. This will draw an equivalent amount of air through the sodium tetrachloromercurate solution in the test tube. Continue with step 3 of the procedure for Environmental Experiment 1-4.

2. Find out more about sulfur dioxide's effects on vegetation. How can you recognize such effects in your area if they exist?
3. Measure the acidity of rain in your area. See Environmental Experiment 2-1 for a suggested technique. How does the pH of the rain compare with the expected norm of 5.7?

ENVIRONMENTAL EXPERIMENT 1-5 • DETERMINATION OF THE EFFECT OF AIR POLLUTION ON MATERIALS

Purpose

To determine the effect of air pollution on materials.

Materials

- | | |
|----------------------|---------------------------|
| 1 pair of scissors | glue, water soluble |
| 1 new nylon stocking | 1 envelope |
| 3 embroidery hoops | masking tape |
| | 30 cm wire, medium weight |

Discussion

Much of the concern with air quality centers on the possible effects of air pollutants on people's health. In addition, damage to vegetation is a serious aesthetic and economic problem as is damage to materials of all kinds. It is well documented that the usable life of many materials is shortened by air pollution. The effects of acid-forming gases such as sulfur dioxide were discussed in Environmental Experiment 1-4. Other acid-forming gases added to the atmosphere by combustion and by industrial and natural processes include carbon dioxide, nitrogen oxides, hydrogen sulfide, and chlorine.

Another damaging agent in the lower atmosphere is ozone which, due to its action as a vigorous oxidizing agent, attacks rubber and textiles. Ozone is an irritating gas to human beings. It affects the eyes, nose, and throat and interferes with normal functioning of the lungs. It is also one of the primary gases responsible for damage to vegetation.

Hydrocarbons in the air are also responsible for damage to materials, especially painted surfaces for which the hydrocarbons act as solvents. Reactive hydrocarbons in the atmosphere are one essential ingredient for the formation of photochemical smog. Smog adds many other irritating and oxidizing chemicals to the atmosphere, creating even more ill effects on human beings, vegetation, and materials.

In this experiment, the focus is not on the analysis of the atmosphere for each of these components, but rather on testing the total effect of polluted air on a representative fabric, nylon.

Procedure

1. Cut 3 pieces of nylon from the nylon stocking. Make each piece larger than the outside dimensions of an embroidery hoop. Stretch a piece of nylon gently but firmly over a hoop and glue it in place. Snap the other half of the mount over the piece of nylon. Trim off the nylon projecting beyond the hoop.

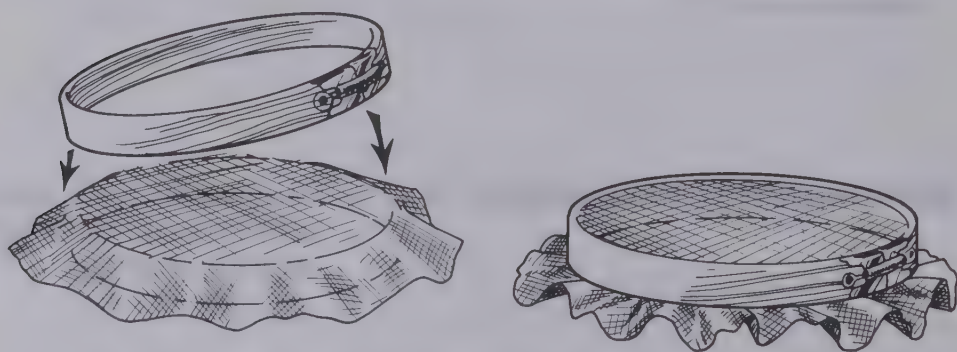


FIGURE 9 One way to prepare pieces of nylon for exposure to sources of air pollutants. Use a small amount of water soluble glue along the rim unless the fit is extremely tight.

Repeat the procedure until you have prepared three identical samples of nylon. (*Hint:* You will find it easier to prepare the samples if you work with a classmate.) Label each sample with a different number.

2. Examine each sample of nylon for any irregularities in the material and for broken threads. Record your findings. Record, too, the color, odor, texture, and elasticity of each sample as well as any other property or properties you may observe.

3. Select one sample of nylon to use as a control. Record the number with which you labeled this sample. Place the sample in an envelope and seal the envelope. Keep the envelope in a safe place.
4. Select two different places where you can expose each of the remaining samples to air. Choose places that are likely to have a high concentration of air pollutants. (Exposing nylon to clean air for a month or more is not likely to result in any obvious changes in the nylon.) Do not select a place where the temperature is high enough to melt or burn the hoop and enclosed nylon. To avoid misleading color changes—changes caused by light—do not select a place where the sample of nylon would be in direct sunlight.

One possible place is the underside of an automobile where the nylon would be exposed to exhaust gases. A fume hood in your kitchen or chemistry laboratory is another good location. Near the end of a runway at an airport, near a busy intersection, near an industrial smokestack, near a smokestack at a power plant, and in a bus terminal are other possible spots to put a sample of nylon.

5. Label each sample with your name and telephone number. Firmly tape or wire each of the samples in a different one of the two places you selected. Record the number on each hoop and its location. Record the date.
6. After a period of 1 month or longer, bring the exposed samples back to the laboratory. Examine each sample for the same properties you observed and recorded in step 2. Record your observations.

Remove any particulate matter from the exposed samples. Compare each of the exposed samples with the control on the basis of each of the properties recorded in step 2. Record your findings.

Questions

1. What effects of air pollutants did you observe in each exposed nylon sample?
2. Suggest reasonable sources of air pollutants at the location of each of your samples of nylon.
3. Answer this question on the basis of your experience in determining the effects of air pollution on materials: Why is it difficult to decide on the dollar value of the economic loss caused by air pollution?

Suggestions for Further Study

1. Rubber also shows marked changes when it is exposed to polluted air. The effects of air pollutants on rubber can be seen more readily if the rubber is stretched during the time it is exposed to the pollutants. Design an experiment to investigate the effects of polluted air on rubber. Use wide rubber (not plastic) bands or strips of rubber cut from a rubber sheet.
2. Are there any public buildings in your area that are showing the effects of air pollution? If there are, what procedures are followed for cleaning the surfaces of these buildings, and how often are the buildings cleaned? Contact officials in charge to learn the answers to these questions.
3. Aging of oil paints is an example of a process that can be greatly accelerated by polluted air. Read and report on the restoration of oil paintings damaged by air pollutants. Someone in a nearby art museum may be able to help you obtain information for your report.

ENVIRONMENTAL EXPERIMENT 1-6 • EXAMINATION OF WASTE PRODUCTS FROM BURNING PLASTICS

Purpose

To examine the waste products from burning plastics.

Materials

1 glass funnel	Bunsen burner
glass wool	matches
1 piece of rubber tubing, 3 cm long	10 test tubes
1 piece of rubber tubing, thick walled, 30-60 cm long	4 small pieces of zinc or magnesium
2 ringstands	4 small chunks of calcium carbonate
1 burette clamp	10 cc dilute hydrochloric acid
1 filter flask or Erlenmeyer flask with two-hole stopper and right-angle glass bend	10 cc silver nitrate solution
distilled water	10 cc sodium cyanide solution
1 10-cc graduated cylinder	15 cc ferrous ammonium sulfate solution
1 right-angle glass bend	pH test paper
1 faucet aspirator or vacuum pump	1 stirring rod
4 samples of plastic taken from such things as a Styrofoam coffee cup, plastic egg carton, plastic tray for meat, plastic bottle, plastic bag, and synthetic cloth	1 water bath
	10 cc concentrated sulfuric acid
	2 cc potassium fluoride solution

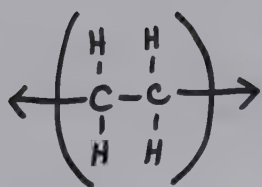
Discussion

Both plants and animals can put together many small chemical units to produce giant molecules known as macromolecules or polymers. Natural cellulose, proteins, silk, and natural rubber are all examples of polymers. If polymers are made synthetically, they may be referred to as plastics. Chemical research of the last 20 years has added significantly to the chemist's understanding of natural polymers. It has also led to the production of a vast array of synthetic polymers or plastics.

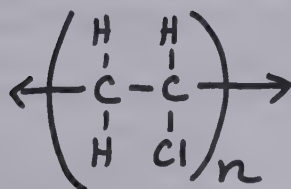
You are familiar with many types of plastics encountered in your everyday life. In fact, the present era is sometimes referred to as the age of plastic. You accept as commonplace plastic wrappings, plastic containers, plastic furniture, and a host of other plastic objects. Some of the more common plastics are nylon, Orlon, Dacron, polyesters, polyvinyl chloride (PVC), polyethylene, Teflon, Styrofoam, Lucite, and Naugahyde (imitation leather). The general structural formulas for some of these substances are shown in Table 1. In each case, the structural unit is repeated over and over again (n times for the general structural formula) to form the polymer.

As with many new technologies, there have been some problems created along with the benefits. The starting materials for many plastics are hydrocarbons derived from petroleum. Therefore, the widespread use of plastics has created another drain on the earth's limited supply of petroleum. In addition, the improper disposal of plastics may create an unsightly mess and thus contribute to the general problem of litter, for plastics are nonbiodegradable substances. Furthermore, when garbage is disposed of by converting it to landfill, the presence of plastics may slow down the conversion process.

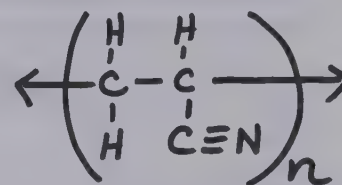
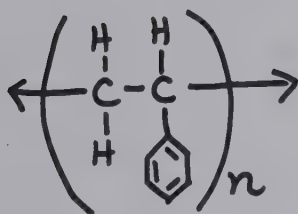
Table 1: General Structural Formulas for Some Common Synthetic Polymers



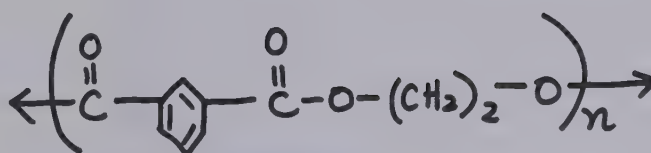
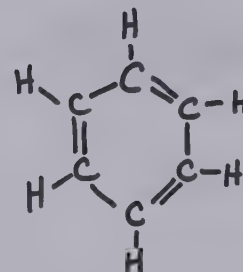
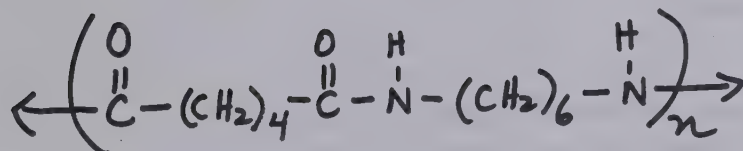
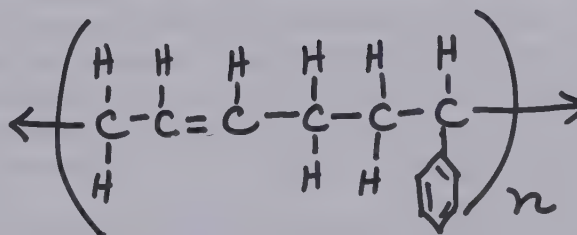
Polyethylene



Polyvinyl chloride (PVC)

Polyacrylonitrile
(Orlon, Acrilan)Polystyrene
(Styrofoam)

stands for

Polyester type of polymer
(Dacron, Mylar)Polyamide type of polymer
(Nylon)Styrene-butadiene rubber (SBR)
(Synthetic rubber)

In this experiment, you will examine another problem that has arisen from the widespread use of plastics. Stated as a question, the problem is: what effect is there on air quality when plastics burn? Possible combustion products of burning plastics include hydrogen chloride, sulfur oxides, organic acids, and a large amount of particulate matter along with carbon dioxide and water vapor. Burning significant amounts of plastics in municipal incinerators may also harm the incinerator itself due to excessive corrosion and clogging of the system by melting plastics.

Procedure

1. Set up the apparatus as shown in Figure 10. Set it in the fume hood if one is available. Although there is little danger of toxic products being given off in high

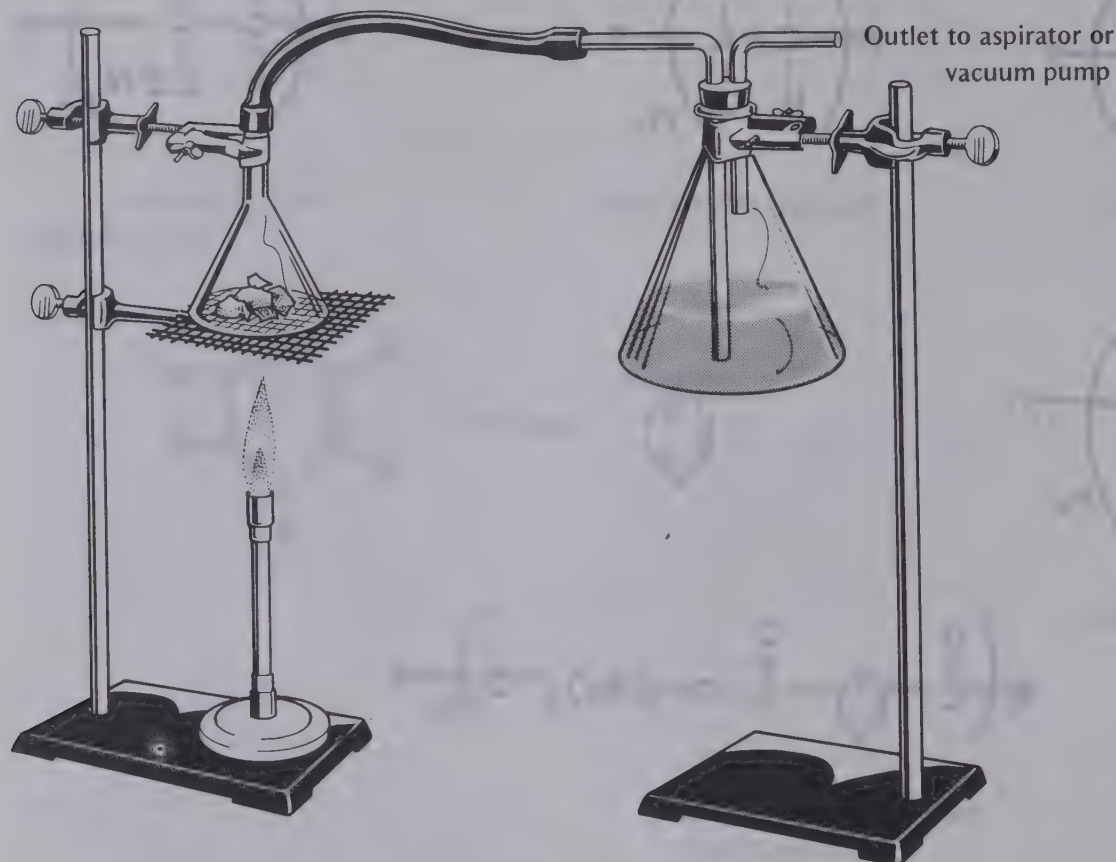


FIGURE 10 Collecting decomposition and combustion products associated with plastics.

concentration when plastic burns, soot given off by the burning of some types of plastics has an annoying way of floating around in the laboratory if it is not confined to the hood.

2. Measure and record the pH of the water in the filter flask.
3. Examine each of the four samples and record their properties.
4. Place the first sample to be burned in the funnel and turn on the faucet aspirator. Ignite the sample by touching the tip of the Bunsen burner flame to it. Observe the burning plastic, watching for the appearance of any gases, particulates, and residue. Record your observations. After the sample has burned, turn off the aspirator.
5. Again measure and record the pH of the liquid in the filter flask.
6. Transfer 5 cc of the liquid in the filter flask to each of 2 test tubes. To one test tube add a small piece of zinc or magnesium. To the other test tube add a small chunk of calcium carbonate. Observe the reaction, if any, in each test tube over a period of 30 minutes. Record your observations.
7. If the gas hydrogen chloride was given off when the plastic burned, it dissolved in the water in the filter flask and dissociated to form hydronium ions and chloride ions. Therefore, a positive test for the chloride ion would indicate that hydrogen chloride gas was given off when the plastic burned.

So you can see what a positive test for the chloride ion looks like, pour 10 cc of dilute hydrochloric acid into a test tube and add silver nitrate solution drop by drop. Record your observations. The precipitate that forms is

silver chloride. The formation of silver chloride upon the addition of silver nitrate to an unknown solution is a positive test for the chloride ion.

To see if the distilled water you added to the filter flask contained the chloride ion, add silver nitrate solution drop by drop to a 10 cc sample of distilled water. Are there chloride ions in the distilled water? Were there chloride ions in the water in the filter flask before you burned the plastic? Record your answer to the last question.

Transfer 10 cc of the liquid in the filter flask to a test tube. Test the liquid for the chloride ion. Were there chloride ions in the water in the filter flask after you burned the plastic? Was hydrogen chloride gas given off when the plastic burned? Record your answers to these questions.

8. If hydrogen cyanide gas was given off when the plastic burned, it dissolved in the water in the filter flask and dissociated to form hydronium ions and cyanide ions. Therefore a positive test for the cyanide ion would indicate that hydrogen cyanide was given off when the plastic burned.

So you can see what a positive test for the cyanide ion looks like, pour 10 cc of sodium cyanide solution into a test tube. CAUTION: Do not inhale fumes from the solution and do not take any of the solution into your body. Add 2 cc of ferrous ammonium sulfate solution and 2 cc of potassium fluoride solution to the test tube. Stir the solutions in the test tube. Then put the test tube into a warm water bath for 2-3 min. With the test tube still in the water bath and stirring constantly, cautiously add concentrated sulfuric acid drop by drop to the test tube. The blue precipitate that forms is ferriferrocyanide. The formation of ferriferrocyanide under the conditions just described when an unknown solution is used instead of sodium cyanide is a positive test for the cyanide ion.

To see if the distilled water you added to the filter flask contained the cyanide ion, repeat the test using 10 cc of distilled water instead of sodium cyanide solution. Are there cyanide ions in the distilled water? Were there cyanide ions in the water in the filter flask before you burned the plastic? Record your answer to the last question.

Transfer 10 cc of the liquid in the filter flask to a test tube. Test the liquid for the cyanide ion. Were there cyanide ions in the water in the filter flask after you burned the plastic? Was hydrogen cyanide gas given off when you burned the plastic? Record your answers to the questions.

9. Discard any residue left in the funnel and any liquid left in the filter flask. Wash and dry the funnel and the filter flask. Assemble the apparatus again, adding 50 cc of distilled water to the filter flask and putting another sample of plastic under the funnel. Repeat steps 3 through 8.

Repeat step 9 two times. Each time use a different sample of plastic.

Questions

1. Use your experimental evidence to discuss the chemical nature of the waste materials produced when different types of plastics burn.
2. What other information would you need to decide whether the incineration of plastics is the best method to dispose of them in your area?

Suggestions for Further Study

1. When plastics were first developed, scientists searched for and found chemical stabilizers to stop or retard reactions caused by light, heat, and bacteria--reactions that broke down, or degraded, the plastics. Ironically, today scientists are trying to modify the formulas of plastics so they will degrade in predictable ways

when exposed to sunlight, heat, and bacteria. Read and report on research on degradable plastics.

2. For some purposes, plastics have largely replaced more traditional materials such as wood, paper, and natural fibers. What types of waste products result when these traditional materials are burned? Design and carry out an experiment to answer this question.

ENVIRONMENTAL EXPERIMENT 1-7 • PURIFICATION OF AIR BY FILTRATION

Purpose

To determine the relative effectiveness of a variety of filters for removing particulate matter from air.

Materials

- | | |
|--|---|
| 4 membrane filters | 4 ring stands |
| 1 balance, sensitive to a milligram | 4 burette clamps |
| 4 identical cigarettes, without filter | 2 filter holders |
| cigarette holder | 1 flow meter (0-5 liters/min) |
| 1 T tube | 1 faucet aspirator or vacuum pump |
| 1 adjustable clamp | matches |
| 1 piece of rubber tubing, 5 cm long | 2 small pieces of different kinds of |
| 3 pieces of rubber tubing, thick walled, 10-20 cm long | material, such as filter paper, cleansing |
| 1 piece of rubber tubing, thick walled, 30-60 cm long | tissue, cotton cloth, and linen cloth |

Discussion

Solid particulate matter emitted by industry is easily visible in the atmosphere.

Therefore, early in the search for methods to control air pollution, technical attention was given to the development of processes for removing particulate matter. As a result, many processes exist today for removing particulate matter from the exhaust gases in which it is suspended.

To pick the best process for removing particulate matter from exhaust gases, one must consider the characteristics of the particles to be removed, including their size and density, and the characteristics of the gases in which the particles are suspended. Of course, he or she must also consider the cost of installing the process and of operating it.

One of the available processes for removing solid particulate matter is filtration. Even though a variety of mechanisms can be used as filters, any filter acts something like a sieve as it strains the particles. The pores of the filter, just like the holes in a sieve, should be smaller than the particles to be removed. The effectiveness of a filter is affected greatly by such factors as the size of the particles, the velocity of the particles, and the type of material used in making the filter. Further, some of the smaller particles in a gas stream which could pass through the pores are also stopped as their forward motion causes some percentage of them to hit the physical parts of the filter. This process is called impaction. Random motion of finely divided particles may also bring several small particles close enough together to increase filtration by impaction.

In this experiment, you will determine the relative effectiveness of filters made of a variety of materials. You will use cigarette smoke as the exhaust gases carrying solid particles. Most of the particles in cigarette smoke consist of unburned tars and are less than $5 \times 10^{-7} \text{ m}$ in size. You can extend your study of filters by carrying out items 1 and 2 in the section entitled *Suggestions for Further Study*. You can use the information you gain from your study in deciding the best procedure to follow for filtering out small particles from other exhaust gases.

Procedure

1. Designate a membrane filter as filter *A* and another as filter *B*. Measure and record the mass of filter *A* and the mass of filter *B* to the nearest milligram.
2. Assemble the apparatus as shown in Figure 11. Be certain to put filter *A* in filter holder *A* and filter *B* in filter holder *B*.

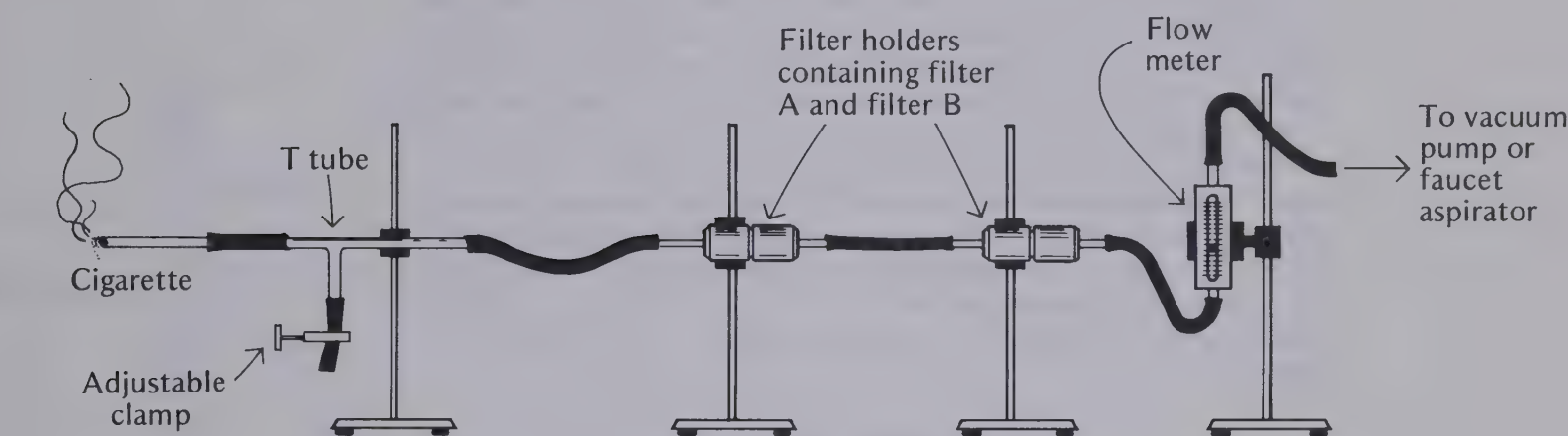


FIGURE 11 Apparatus for collecting smoke particles on filters.

3. To gain experience operating the apparatus, turn on the faucet aspirator or vacuum pump without lighting the cigarette. Adjust the flow rate to 1 liter/min. At this flow rate, a cigarette should burn in approximately 5 min.
4. With the faucet aspirator or vacuum pump turned on and the flow rate adjusted to 1 liter/min, light the cigarette. Record the exact conditions under which the cigarette burns.
5. When the entire cigarette has burned, turn off the faucet aspirator or vacuum pump. Examine the apparatus for particles not removed by the filters. Record your findings, if any. Remove each filter and again measure and record its mass.
6. Cut 2 filters of different materials to fit filter holder *A*. Measure and record the mass of each filter. Measure and record the mass of each of two new membrane filters. Repeat steps 4 and 5 twice. Each time use a different one of the filters you cut for use in filter holder *A* and a different one of the new membrane filters in filter holder *B*.

Questions

1. Calculate the mass of particulate matter trapped on each filter (final mass of filter—initial mass of filter).
2. Calculate the percent of the total mass of particulate matter trapped on filter *A* for

each type of material used. Assume that all particulate matter is trapped on one or the other filter and then apply this formula.

$$\text{percent of particulate matter on filter A} = \frac{\text{mass of particulate matter on filter A}}{\text{mass of particulate matter on filter A} + \text{mass of particulate matter on filter B}} \times 100$$

3. On the basis of your calculations in Question 2 and your observations, discuss the relative effectiveness of the different types of filters you tested.
4. How could the information you collected on the effectiveness of different materials in removing solid particulate matter be applied in industry?

Suggestions for Further Study

1. The apparatus assembled for this experiment could also be used to study how variables other than the composition of the filter affect the effectiveness of the filtering process. For example, what effect, if any, does an increase in the flow rate of the cigarette smoke have on the percent of the total mass of particulate matter trapped on filter A? Design and carry out an experiment to answer this question.
2. Do different brands of cigarettes give off different total amounts of particulate matter when they burn? Design and carry out an experiment to answer this question. Compare your results with the tar content given in a published list for the brands of cigarettes you used.
3. How is particulate matter from automobile emissions controlled? Read and report on this subject.

ENVIRONMENTAL EXPERIMENT 1-8 • PURIFICATION OF AIR BY ELECTROSTATIC PRECIPITATION

Purpose

To demonstrate the process of electrostatic precipitation for the removal of particulate matter from air.

Materials

- | | |
|--|--|
| 1 faucet aspirator or vacuum pump | 1 spark coil |
| rubber tubing, thick walled, 30-60 cm long | copper wire, insulated, 60-100 cm long |
| cigarette | 1 power supply (0-7 volts) |
| 1 condenser | matches |
| 2 meters copper wire, uninsulated, fine | 5 cc acetone |
| 1 ring stand | 1 evaporating dish |
| 2 burette clamps | |

Discussion

The need to improve the quality of the air and the need to recover valuable particulate matter motivate industries to remove solid particles from exhaust gases. A device that can be used to fulfill either or both of these needs is an electrostatic precipitator. An

electrostatic precipitator operates on the principle that unlike electric charges are attracted to each other. The first electrostatic precipitator was developed by Frederick G. Cottrell in 1910, and, therefore, the device is often called the Cottrell precipitator.

Suspended particles adsorb electric charge on their surfaces. Particles of the same material adsorb the same kind of electric charge. Because like electric charges repel each other, the particles remain suspended in the exhaust gases. Therefore, the collecting surface in an electrostatic precipitator is given an electric charge opposite to that on the particles to be removed.

As the stream of exhaust gases flow past the collecting surface, the charged particles in the exhaust gases are attracted to the oppositely charged collecting surface. Once the charged particles hit the collecting surface, their charges are neutralized and the particles settle out. The collected particles can be recovered by mechanical means such as scraping the collecting surface or by washing down the collecting surface with an appropriate liquid. Modern electrostatic precipitators are almost 100% effective in removing suspended particles from exhaust gases when the diameters of the particles fall within the range of 10^{-7} – 10^{-4} m.

Procedure

1. Set up the apparatus as shown in Figure 12. Do not plug in the power supply and do not light the cigarette. Be sure no bare wire is touching the ring stand or a burette clamp. Ask your teacher to check and approve the apparatus.

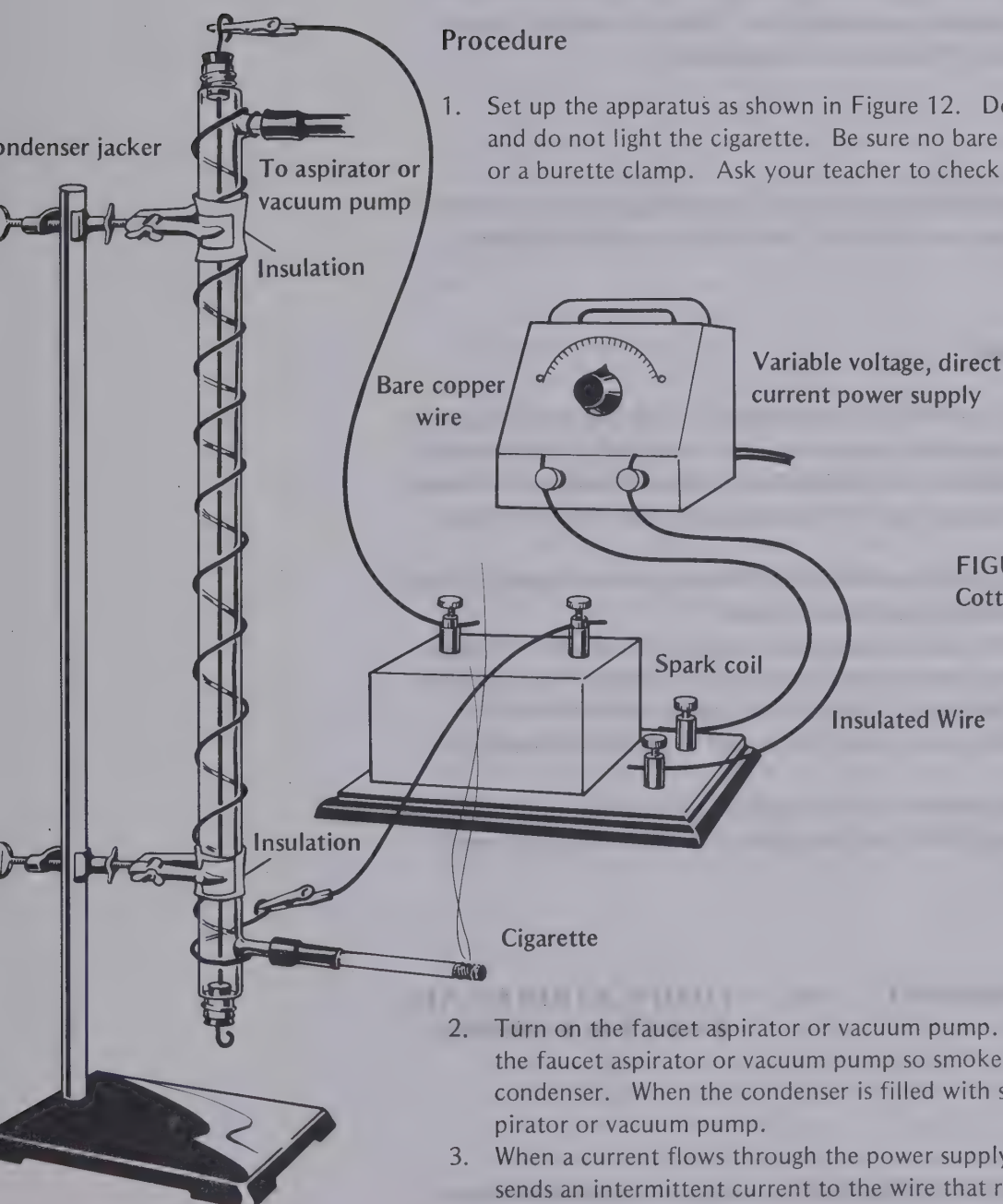


FIGURE 12 A demonstration Cottrell Precipitator.

2. Turn on the faucet aspirator or vacuum pump. Light the cigarette. Adjust the faucet aspirator or vacuum pump so smoke is drawn slowly through the condenser. When the condenser is filled with smoke, turn off the faucet aspirator or vacuum pump.
3. When a current flows through the power supply, the vibrator on the spark coil sends an intermittent current to the wire that runs through the inside of the condenser. CAUTION: Do not touch or get close to any of the electrical

terminals or to the uninsulated wire when the power supply is turned on. You could get a shock.

Plug in the power supply and slowly increase the voltage until you see some effect on the smoke accumulated in the condenser. If you should send a high enough voltage through the wire, you will see a spark jump between the wire inside the condenser and the wire coiled around the outside of the condenser. Should this happen before you observe an effect on the smoke, do not continue to increase the voltage. Instead adjust the voltage just below the sparking point. Record your observations. Turn the voltage down to zero and unplug the power supply.

4. Repeat steps 2 and 3 two times.
5. With the power supply unplugged, disconnect the wires from the condenser to the spark coil. Remove the partially burned cigarette from the condenser and the rubber tubing to the faucet aspirator. Remove the condenser from the ring stand.

Add 5 cc of acetone through one of the side arms of the condenser. Rotate the condenser to rinse the collected particulate matter off the inside of the condenser. Pour the acetone into an evaporating dish. Allow the acetone to evaporate and examine the residue. Record your observations.

Questions

1. How effective was your electrostatic precipitator in removing particulate matter?
2. What possible limitations can you think of to the routine use of electrostatic precipitation in industry?

Suggestions for Further Study

1. If a spark passes through air, some oxygen molecules, O_2 , are converted to ozone molecules, O_3 . Lightning causes this reaction to occur naturally; it can also take place in electrostatic precipitators. Is it desirable to produce ozone as a by-product of particulate removal? Read and report on the properties and effects of ozone in the lower atmosphere.
2. If particulate matter cannot be removed by either filtration or electrostatic precipitation, what alternative processes might then be used?
3. What industry or industries in your area use electrostatic precipitation? Choose one of the industries and find out how long it has used this method for removing particulate matter. Visit, telephone, or write to an appropriate official of the industry you chose to find out more about the control of particulate matter in your area.
4. Who was Frederick Cottrell and how did the invention of the electrostatic precipitator fit into his entire career? What were his other interests? Read and report on this American scientist.

ENVIRONMENTAL EXPERIMENT 1-9 • PURIFICATION OF AIR BY WET SCRUBBING

Purpose

To demonstrate the process of wet scrubbing for the removal of contaminating gases from air.

Materials

- | | |
|---|--------------------------------------|
| 2 ring stands | 1 condenser without inner tube |
| 3 burette clamps | glass wool |
| 2 rings | glass beads or small pieces of glass |
| 2 wire gauzes | tubing to fit condenser |
| 1 Bunsen burner | 4 beakers, 250 cc |
| 1 Florence flask, Pyrex, 500 cc | 5 grams manganese dioxide or 25 |
| 1 two-hole rubber stopper to fit Florence flask | grams marble chips |
| 1 thistle tube | pH paper |
| 2 right-angle glass bends | 6 M hydrochloric acid |
| 1 piece of rubber tubing, 3 cm long | matches |
| 1 piece of rubber tubing, 30-60 cm long | 1 Erlenmeyer flask, 250 cc |
| | 150 cc of 1 molar sodium thiosulfate |

Discussion

Removing contaminating gases from a stream of exhaust gases calls for different strategy than removal of solid particulate matter. Because molecules of gases are not large enough to be removed by filtration or to be acted on by an electrostatic precipitator, other means must be used.

One of the most widely used methods depends on the absorption of contaminating gases by water or another suitable liquid. To make sure the gases have a good probability of being absorbed in the water, it is necessary to provide a very large surface area between the water and the gases. A large surface area gives the gases many opportunities to come in contact with the water. On an industrial scale, a large surface area is provided by filling a tower with many pieces of inert material. Ceramic and glass materials are used.

The stream of exhaust gases enters the tower at the bottom and rises through water flowing down over the pieces of inert material. As the exhaust gases flow upward through the tower, the soluble contaminating gases in the exhaust gases have many opportunities to dissolve in the water. Eventually, partially or completely cleaned, or scrubbed, exhaust gases emerge from the top of the tower. The water with the soluble contaminating gases dissolved in it (the water solution of gases) drains out the bottom of the tower. This process for removing contaminating gases from exhaust gases is called wet scrubbing. It is carried out in a wet scrubber which is shown in Figure 13.

What gases can be successfully removed by the process of wet scrubbing? Those that are soluble in water or react with water are most likely to be removed. Some common gaseous contaminants that are removed by wet scrubbing are sulfur dioxide, hydrogen sulfide, chlorine, nitrogen oxides, and hydrogen fluoride. Contaminating gases which are soluble in a solvent other than water can be removed by passing them through a tower in which the solvent is used instead of water.

Procedure

1. Chlorine is a good gas to use as the contaminating gas. You can follow its progress through the scrubber because chlorine is colored and quite soluble. **However, chlorine is a toxic gas and can be harmful to you if it is not prepared in a hood or well ventilated place.** If a suitable place for generating chlorine is not available, carbon dioxide can be used instead of chlorine as the contaminating gas. Consult your teacher as to which gas you should use.
2. Set up the apparatus as shown in Figure 13. If you are going to generate



FIGURE 13 A demonstration wet scrubber and gas generator.

- chlorine, be certain to set up the apparatus in the hood or in a well ventilated place. Ask your teacher to check and approve your apparatus.
3. Remove the stopper from the Florence flask. If you are generating chlorine, add manganese dioxide to the flask. If you are generating carbon dioxide, add chunks of calcium carbonate to the flask. Replace the stopper.
 4. Turn on the faucet so the water just trickles through the stopper. The water should run slowly enough for it to flow down between the small pieces of glass tubing packed in the scrubber. Test and record the pH of the water as soon as it flows out the bottom of the scrubber.
 5. To start the generator, add enough dilute hydrochloric acid through the thistle tube to cover the bottom of the thistle tube. This will prevent gas from escaping up the thistle tube. If the generator stops producing gas, add more acid. If you are producing chlorine, gently heat the reaction mixture. If you are producing carbon dioxide, you need not heat the reaction mixture.
 6. Observe the gas as it enters and rises through the scrubber. Record all of your observations. Test and record the pH of the gas escaping out the top of the scrubber by holding a piece of wet pH paper at the outlet.
 7. After approximately 200 cc of water have collected in the beaker at the bottom of the scrubber, replace the beaker with another. Repeat this procedure until you have collected 3 or 4 samples of water. Number the beakers in the order in which you collect the samples. Test and record the pH of each sample.



FIGURE 14 Arrangement for getting rid of surplus chlorine gas. As the generator flask cools, pressure in it will go down. Lift the glass tube in the sodium thiosulfate to just above the solution so that it will not back up into the generator.

8. To stop generating chlorine, stop heating the flask. Study Figure 14 so you will know exactly what to do with the excess chlorine in the generator. Then disconnect the side arm of the flask where it is attached to the scrubber. Immediately attach a glass bend to the side arm. Put the other end of the glass bend in 1 molar sodium thiosulfate solution in a flask. Sodium thiosulfate solution will react with the excess chlorine and prevent it from escaping into the room.

To stop generating carbon dioxide, stop adding acid. You can add water slowly through the thistle tube to dilute the acid in the generator. Disconnect the side arm from the scrubber and allow excess carbon dioxide to escape into the air in the room.

Questions

1. How effective was your wet scrubber in removing the gas? Which one of your observations supports your answer?
2. Would particulate matter be removed by a wet scrubber? What types of contaminants would not be removed by this device?
3. What effect did the contaminating gas have on the water flowing through the scrubber? What problem, if any, would this effect create in industrial uses of the process of wet scrubbing?
4. Write the chemical equation for the reaction that took place in the chlorine generator. In the carbon dioxide generator.

Suggestions for Further Study

1. Visit, telephone, or write to officials at power plants and industries in your area to find out if wet scrubbing is used for the removal of gaseous contaminants.

- Report your findings. Are there any industries in your area that could use wet scrubbing to reduce or eliminate contaminants but instead use an alternate method?
2. Sulfur dioxide has been such a prevalent air pollutant that much of the research on removal of gaseous contaminants has been directed at the removal of sulfur dioxide. Wet scrubbers are one method commonly used to remove sulfur dioxide. Read and report on other methods used to remove sulfur dioxide from exhaust gases.
 3. Instead of using liquids such as water to absorb contaminating gases, purifiers can use such materials as activated charcoal, silica gel, and activated alumina to adsorb contaminating gases. What are these substances and why are they useful in removing gases? What are the advantages and the disadvantages of using the process of adsorption to remove contaminating gases instead of the process of absorption?
 4. Just as gaseous contaminants added to the environment can be harmful, so can heat added to the environment be harmful. As you have learned, gaseous contaminants can be removed by wet scrubbers. Heat can be removed by cooling towers. Read and report on the design and operation of one or more different kinds of cooling towers.

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Brodine, Virginia, *Environmental Issues: Air Pollution*, New York: Harcourt Brace Jovanovich, 1973.

This book surveys the entire air pollution question. It is a well written book for the general reader. The author is the consulting editor of *Environment* and has written many articles about air pollution for that magazine. (Some of the material in this book appeared in Jan/Feb and May, 1972, issues of that journal.) Has there been any progress in combating air pollution? What are the health effects and the environmental effects? How are standards set? These, and many other questions, are considered in this book.

Environmental Protection Agency, Air Pollution Control Office, *Air Quality Criteria for Carbon Monoxide*, 1970. *Air Quality Criteria for Hydrocarbons*, 1970. *Air Quality Criteria for Nitrogen Oxides*, 1971. *Air Quality Criteria for Particulates*, 1969. *Air Quality Criteria for Photochemical Oxidants*, 1970. *Air Quality Criteria for Sulfur Oxides*, 1967. Washington, D.C.: U.S. Government Printing Office. Costs range from \$1.50 to \$2.00.

This series of publications details what scientists have been able to determine so far about the effects of each category of air pollutant. The Environmental Protection Agency (EPA) then uses these statements of observed scientific knowledge as criteria, to develop standards. EPA standards are legal limits of pollution based not only on the criteria but also on judgments of benefit vs. risk, economic costs, and other societal determinations.

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One of the main determinants of air quality in a region is the method chosen to generate electricity in that region. This text, which was developed for secondary schools and classes in adult education, examines the impacts of generating electricity with fossil fuels such as coal and oil versus the impacts of generating electricity with nuclear fuel. For each method of generating electricity, the impacts on air quality are discussed along with the problems of thermal effects and disposal of waste products. The book has a long glossary containing terms related to power production by fossil and nuclear fuels.

Pryde, Lucy T., *Chemistry of the Air Environment*, Menlo Park, California: Cummings Publishing Company, 1973. (80 pages, softcover)

This booklet will give you more information about air quality. It discusses the sources and types of air emissions and their chemical properties, expected effects, and means of control. Special attention is given to the role of the automobile as a polluter. A glossary of environmental and chemical terms is included.

Science Magazine, April 19, 1974.

This issue is devoted to articles on energy policy, economics, and alternative sources of energy. It includes information on the developing technology of solar and geothermal energy.

PART TWO INVESTIGATING WATER QUALITY

OBJECTIVES

By the time you have completed Part Two of this module, you will be expected to demonstrate acceptable performance on the following learning objectives:

1. Describe the *hydrological cycle*.
2. Explain why the quality of surface and ground waters is important to you.
3. Measure the pH of samples of water and discuss the importance of the pH measurement in determining the quality of water.
4. Define the term *hard water* and explain its significance.
5. Measure the concentration of dissolved solids in water.
6. Determine the amount of dissolved oxygen in samples of water and explain the meaning of each measurement.
7. Measure the concentration of nitrate ion in samples of water and describe the role of nitrate ions in water quality.
8. Determine the concentration of phosphate ion in samples of water and explain the relationship between the concentration of phosphate ion and water quality.
9. Define the word *eutrophication* and discuss the role of nitrate ions and phosphate ions in bringing about rapid eutrophication of a lake.
10. Test for the presence of the lead ion in water, identify the source of this ion, and discuss its impact on people's health.
11. Determine the presence of coliform bacteria in water and explain the significance of this presence in water.
12. Summarize each of the tests you have performed to determine the quality of water.
13. Explain the process of coagulation and discuss its usefulness in purifying water.
14. Discuss several different types of filters and explain how filtration can be used in purifying water.
15. Explain the process of distillation and the purpose it serves in the purification of water.
16. Explain the process of ion exchange and discuss its application to the purification of water.
17. Summarize each of the methods you have tested for purifying water.

SUGGESTED ORDER OF STUDY

1. Review the physical and chemical properties of water, Section 3-5 in *Keys to Chemistry*.
2. Study Environmental Sections 2-1 and 2-2.
3. Environmental Experiments 2-1 to 2-7 deal with the analysis of water. Read through these experiments and decide which ones you wish to carry out. Do the experiments that you choose to do in numerical order. Hand in a report for each experiment.
4. Environmental Experiments 2-8 to 2-11 deal with the purification of water. Read them and decide which ones you wish to carry out. Again, do the ones you chose in numerical order. Hand in a report for each experiment.
5. Explore the *Bibliography* for Part Two for sources of additional information on the analysis and purification of water.
6. Check with your teacher for evaluation of your learning experience with Part Two of this module.

ENVIRONMENTAL SECTION 2-1 • INVESTIGATING THE HYDROSPHERE

All the water on the surface of the earth and all the moisture in the atmosphere surrounding the earth make up the *hydrosphere*. Water is tremendously important to living things. Plants and animals use water to carry on their life activities. Homes and industries use water, too. Furthermore, water provides people with recreational activities such as swimming, fishing, and boating.

Water, a very versatile and important chemical, is found in enormous quantities on the earth. Approximately 75 percent of the earth's surface is covered with water. However, more than 97 percent of the available water is found in salty oceans, quite unsuitable for many purposes. If salt could be removed from ocean water inexpensively, the oceans would become a vast source of water suitable for human needs. In contrast, less than 1 percent of the earth's entire water supply is found in surface and ground waters, the usual sources of water used by people. Since surface water--lakes and rivers--and ground water are in relatively short supply, it is important that the quality of these waters not be impaired.

Air and water both move in large scale, dynamic circulation cycles. Look at Figure 15. It is a simplified diagram of the hydrological cycle, the pattern in which

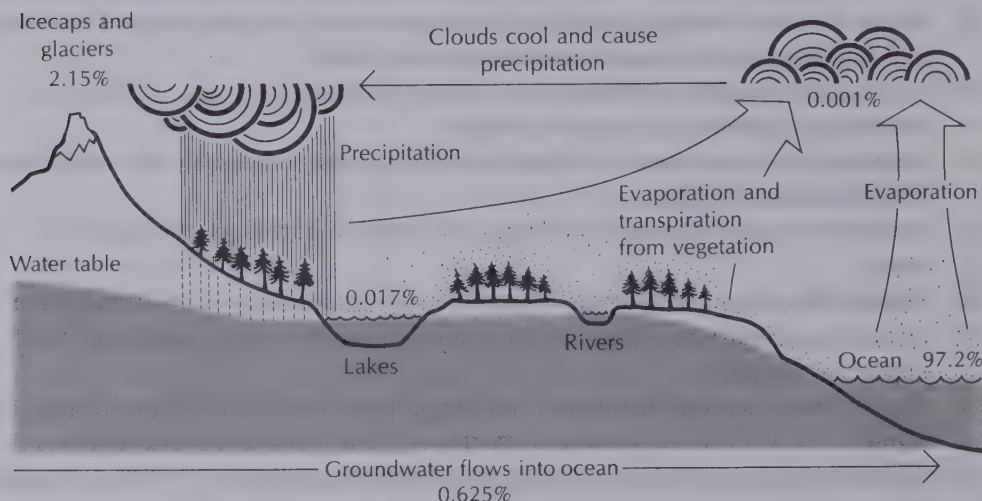


FIGURE 15 The Hydrological Cycle

water circulates over and over again. Water evaporates from the ocean into the atmosphere, changing from a liquid to a vapor. As the atmosphere circulates, the water vapor is carried over the land. Water from vegetation, rivers, and lakes also passes into the atmosphere as water vapor. When water vapor cools sufficiently, it forms clouds from which precipitation falls. The flow of ground water and rivers carries the water back into the ocean where it will evaporate again to make the hydrological cycle continuous.

Actually, the hydrological cycle also includes many smaller patterns of circulation instead of one large pattern as shown in Figure 15. For example, water that evaporates from the ocean today may return to the ocean tomorrow in the form of precipitation. The hydrological cycle is a natural method for recycling and purifying water.

This part of the module lets you investigate ways to test the quality of water and various methods for purifying water.

ENVIRONMENTAL SECTION 2-2 • GUIDELINES FOR COLLECTION OF WATER SAMPLES

For the results of the tests you perform on your samples of water to be meaningful, you will need to use good procedures for collecting your samples. Therefore, as you gather your samples of water, keep these points in mind:

Sampling

Before you leave the laboratory, wash and rinse the bottles you will use to hold the samples of water you collect. In the field, rinse each bottle two or three times with the water to be collected in it. Then fill the bottle with a sample of water. Remember that each sample you collect should be representative of its source. Avoid collecting suspended matter unless it really is characteristic of the body of water being sampled. If you are sampling tap water or well water, let the water run awhile before collecting a sample. When sampling large bodies of water, record the depth at which you collect each sample.

Do not attempt to obtain samples of water containing untreated sewage or untreated industrial wastes without securing permission from officials in charge and without taking measures to protect yourself. It would be safer for you to obtain samples of sewage and industrial wastes after they have been treated and discharged into receiving waters. However, it would still be wise for you to protect yourself by wearing rubber or plastic gloves.

Timing

As a general rule, the shorter the time between collection and analysis of a sample of water the more accurate your results will be. Some measurements are more likely than others to be affected when the samples have been stored before being analyzed. For this reason, measurement of temperature, amount of dissolved oxygen, and pH are usually carried out in the field immediately after the sample is collected. For determining most other measurements, 12 hours is generally considered to be the maximum time limit between collection and analysis for samples of polluted waters, and up to 72 hours as the maximum time limit for storing samples of unpolluted water.

Changes in water caused by bacterial action can be minimized by storing the samples in a cold, dark, place. Furthermore, samples that will not be tested for

bacterial content can be stabilized by the addition of chloroform or other germicide (1 cc/100 cc of sample). The action of chlorine in samples of treated waters can be stopped by the addition of sodium thiosulfate (10 mg/100 cc of sample).

Recording

Be prepared to label all samples at the time you collect them. Waterproof felt pens or labels are helpful for this purpose. Don't trust your memory, as it is amazing how much alike samples of different water look once you are back in the laboratory. At the time you collect each sample, record in your laboratory notebook the source of the sample, the date and time you collected it, and any other pertinent information.

ENVIRONMENTAL EXPERIMENT 2-1 • MEASUREMENT OF pH

Purpose

To measure the pH of different samples of water.

Materials

clean sample bottle for each sample to be tested	selection of chemical indicators such as solutions of litmus, phenolphthalein, bromthymol blue, and universal indicator as solution or paper
samples of water	glass stirring rods
1 spotplate, individual watch glasses or evaporating dishes, or plastic wrap over white paper	

Discussion

Chemical indicators are used to tell if a water solution is acidic, basic (alkaline), or neutral. This simple determination is made regularly on samples of drinking water as well as other samples of water for which the acid-base character of the water must be known. Water in natural systems such as lakes or rivers normally has a pH between 5 and 9, but the pH of the water can be changed by the intentional or unintentional addition of many substances.

Generally one step in the purification of drinking water is the adjustment of the pH of the water to slightly alkaline, usually between a pH of 7.0 and 8.5. This adjustment helps to prevent corrosion of waterlines. Often water that has been used in an industrial process or is the effluent from municipal waste must also be adjusted in pH before discharge in order to meet standards for water quality. Highly acidic or highly alkaline waters can be extremely harmful to wildlife and vegetation, although many organisms can tolerate fairly sudden, severe pH changes.

Procedure

1. Select three or more sources of water to sample. You might select different sources of drinking water, or you might select a river, a lake, a small stream, a

puddle, or an ocean. You might be able to obtain samples of water from water treatment plants or various agricultural or industrial operations in your area.

Collect a sample of water from each source you selected. Use a clean bottle to hold each sample. Obtain each sample close to the time you will be measuring the pH of the water. Although you will need only a few drops of water to measure the pH of the sample, you may want to use these same samples for other experiments in Part Two, so plan ahead.



FIGURE 16 Materials for testing pH of water samples.

- With each of the indicators you have available, determine the approximate pH of each of your samples of water. The apparatus is illustrated for you in Figure 16. Compare colors with the color key on page 121 of your *Keys to Chemistry* text or to other color keys that are available to you. Record your results in a table. Table 2 is an example of a record of data. The data is based on a sample of water from the author's kitchen faucet.

You can find more information on acid-base indicators in *The Handbook of Chemistry and Physics*.

Table 2: The pH of a water sample

Water sample	Indicator	pH range in which expected color change takes place	Color of indicator in sample	Approximate pH
Drinking water, kitchen faucet in home	phenolphthalein	8.3 --- 9.8 colorless pink	colorless	≤ 8.3
"	bromthymol blue	6.0 --- 7.6 yellow blue	blue	≥ 7.6
"	universal indicator	continuous	blue-green	≈ 8

3. If your indicator is in solution form, add several drops of it to a small amount of sample in the depression of a spotplate. If the indicator is in paper form, dip a clean stirring rod in the sample and transfer a drop to the indicator paper.

Questions

1. On the basis of your data, estimate the pH for each of your samples of water. (For example, the sample of water for which the data is recorded in Table 2 would be reported as having a pH between 7.6 and 8.3 or approximately 8.)
2. Does the pH of each sample show the water to be suitable for its intended use? State your reasoning for each answer.
3. If the pH indicates a sample of water to be very acidic or very basic, try to find out the chemical cause of the observed result.

Suggestions for Further Study

1. If your supply of drinking water is acidic, what chemical(s) could you use to adjust the pH toward alkalinity? If it is too alkaline, what chemical(s) could you use to adjust the pH toward acidity?
2. What effect does the addition of chlorine to drinking water have on the pH of the water?
3. Water pipes were once commonly made from lead; such pipes have now been largely replaced by galvanized pipes (zinc coated) and copper pipes. Copper pipes are often joined with solder that has a high lead or cadmium content. Asbestos-cement pipe is also commonly used, and plastic pipes made of polypropylene are just coming into use. Which type of pipes are used in your home or school? What effect would acidic water have on each type of pipe?
4. Small lakes in which the pH of the water differs may support different populations of organisms. See if you can find two similar lakes with different pHs, and compare the organisms present. What other factors besides the pH could influence the kinds of organisms in a lake?
5. If your family happens to have a swimming pool or you have worked at a public pool, you may be interested in investigating how pH is adjusted in a pool and for what reasons.

ENVIRONMENTAL EXPERIMENT 2-2 • DETERMINATION OF DISSOLVED SOLIDS

Purpose

To measure the amount of dissolved solids in a sample of water.

Materials

- | | |
|--|---|
| 1 clean sample bottle for each water sample | 1 porcelain evaporating dish or beaker of greater than 100 cc capacity |
| 1 funnel, fine pore filter paper, 1 250-cc beaker, 1 ring stand, 1 ring, and 1 clay triangle (if the sample is cloudy) | distilled water |
| 1 activated charcoal filter (if sample is cloudy and filtration does not remove all the cloudiness) | 1 thermostatically controlled drying oven or 1 ring stand, 1 ring, 1 wire gauze, 1 Bunsen burner, and matches |
| | 1 balance |
| | 1 volumetric flask, 100 cc |

Discussion

What appears to be a sample of pure water is unlikely to contain only water molecules. Water is such a good solvent that most bodies of water contain a complex mixture of dissolved materials as well as suspended materials. One particular set of materials—the dissolved ions of calcium, magnesium, and iron—contribute to the property of water known as hardness. These ions are responsible for the gummy curd formed by the reaction of soap in hard water.

The chemicals that make water hard are the carbonate, bicarbonate, and sulfate compounds of calcium, magnesium, and iron. When hard water is heated, the bicarbonates in the water decompose to form carbonates. As the water evaporates, the sulfates and carbonates in the water form a white residue that clogs hot-water pipes and deposits on the inside of your teakettle. On the industrial level, the white residue is called boiler scale. Boiler scale creates serious problems. Not only does it clog pipes, but it also prevents effective transfer of heat.

Although some of the chemicals dissolved in water are annoyances under some circumstances, dissolved materials are often the source of essential nutrients to plants and animals. Another aspect of these dissolved materials is the toxic effect they can have if the concentration is too high. An example is the over fertilization of a lawn or garden plot which “burns” the plants.

Procedure

1. Obtain a sample of tap water, well water, softened water, rain water, or other water of your choice. You will use 100 cc of the sample in this experiment.
2. If the sample of water is cloudy, filter it using fine pore filter paper. Filter the water several times if the cloudiness persists. If the water is still cloudy after repeated filtration, filter it through activated charcoal. Ask your teacher for help in setting up an activated charcoal filter.
3. Clean, rinse with distilled water, and then wipe dry the procelain evaporating dish. To make certain the evaporating dish is thoroughly dry, put it in the oven set at 100°C. If an oven is not available, warm the evaporating dish over a Bunsen burner. Be sure that you use a well adjusted flame so it doesn't deposit carbon on the evaporating dish. Allow the evaporating dish to cool to room temperature. Measure its mass using the highest precision balance available. Record the mass of the evaporating dish.
4. Use the volumetric flask to measure out 100 cc of your sample. Transfer the 100 cc of water in the volumetric flask to the preweighed evaporating dish. Again measure and record its mass.
5. Place the evaporating dish in the oven or over the Bunsen burner and evaporate the water. If a Bunsen burner is used, heat gently so as to avoid spattering and thus loss of some of the sample. Remove the evaporating dish, which now contains only a dry residue, from the oven and allow it to cool. Then measure and record the mass of the evaporating dish and its contents.

Questions

1. What did you expect the mass of the 100 cc of water to be? What was its mass?
2. What was the mass of the dissolved solids in the sample of water you tested?
3. Calculate the concentration of the dissolved solids in parts per million (ppm).

This is the relationship that will enable you to make the calculation:

$$\frac{\text{grams of dissolved solids}}{\text{grams of water tested}} \times 10^6 = \text{parts per million by mass}$$

4. The U.S. Public Health Service recommends 500 ppm as the maximum concentration of dissolved solids in drinking water. How does the amount of dissolved solids in your sample compare with the federal standard of drinking water?
5. Where did the dissolved solids in your sample of water come from?

Suggestions for Further Study

1. What is the reproducibility of the method used in this experiment to determine the concentration of dissolved solids in water? To find the answer to this question, use the method in this experiment to determine the concentration of dissolved solids in five or more samples of water from the same source. Is the concentration of dissolved solids in each sample the same? If not, find the difference in percent among your results.
2. What effects, if any, does the drinking of very hard water have on people's health? What effects, if any, does the drinking of very soft water have on people's health?
3. Obtain samples of two or more brands of bottled water. Determine the concentration of dissolved solids in each sample. If you were going to buy bottled water to drink, what influence, if any, would your results have on your choice of brand of bottled water? What factors other than concentration of dissolved solids would determine whether or not you would buy bottled water?
4. Determine the relative concentration of calcium ions in two or more samples of water. Use the procedure described in step 3 of Environmental Experiment 2-10.
5. In areas where well water has been used to irrigate crops, the dissolved solids can build up in the soil and render it unsuitable for farming. How do the farmers cope with this problem?

ENVIRONMENTAL EXPERIMENT 2-3 • MEASUREMENT OF DISSOLVED OXYGEN

Purpose

To measure the amount of dissolved oxygen in a sample of water.

Materials

- | | |
|-----------------------------|-----------------------|
| 1 dissolved oxygen sampler | 1 Celsius thermometer |
| 1 dissolved oxygen test kit | |

Discussion

Dissolved oxygen is essential for fish and other aquatic life. The oxygen is also essential for the normal purification processes carried out by microorganisms in the water. For example, when waste matter is placed into a body of water, some or all of the dissolved oxygen is used up by the microorganisms, such as bacteria, as they decompose the matter. In extreme cases, the amount of dissolved oxygen falls to zero. Then the bacteria that use dissolved oxygen (aerobic bacteria) are replaced by bacteria that take the oxygen they need from chemical compounds in the waste matter (anaerobic bacteria). Further decay carried out by the anaerobic bacteria produces the unpleasant and toxic gases associated with rotting and stagnant waste.

The determination of the concentration of dissolved oxygen in a sample of water is probably the best single test that can be performed to indicate overall water quality. By comparing the concentration of dissolved oxygen present in the water to the maximum concentration that could be present, an indication of the purity of the body of water is obtained.

The maximum amount of dissolved oxygen in any body of water depends largely on the temperature of the water. As you know from your experience, as a glass of carbonated liquid warms up, it fizzes. The fizz is bubbles of carbon dioxide that are going from the dissolved condition to the gaseous condition. So it is with water and oxygen; cold water can dissolve considerably more oxygen than can warm water. The following table gives the maximum amount of dissolved oxygen in water at various temperatures.

Table 3 Maximum Concentration of Dissolved Oxygen in Water at Varying Temperatures and Constant Pressure

Temperature °C	Dissolved Oxygen (ppm)	Temperature °C	Dissolved Oxygen (ppm)
0	14.6		
1	14.2	16	10.0
2	13.8	17	9.8
3	13.4	18	9.6
4	13.2	19	9.4
5	12.8	20	9.2
6	12.4	21	9.0
7	12.2	22	8.8
8	11.8	23	8.6
9	11.6	24	8.5
10	11.4	25	8.4
11	11.0	26	8.2
12	10.8	27	8.0
13	10.6	28	7.9
14	10.4	29	7.9
15	10.2	30	7.6

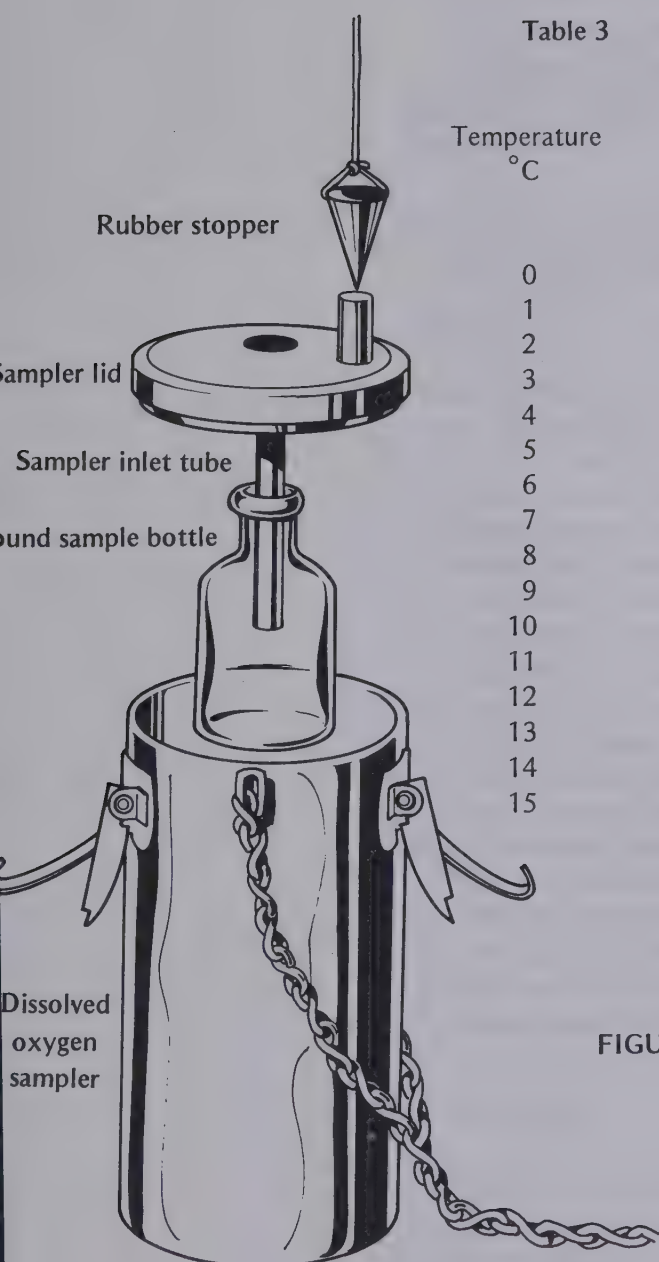


FIGURE 17 The parts of a dissolved oxygen sampler.

Procedure

This is one of a very few experiments in which you will be asked to use a kit. The use of a kit will allow you to make determinations of the concentration of dissolved oxygen in the field. If a sample of water is carried from the field to the laboratory,

its concentration of dissolved oxygen can change. Therefore, for the results of a test to be an accurate measure of the concentration of dissolved oxygen in the body of water sampled, the test must be made in the field at the time the sample is taken.

1. Rinse the round sample bottle, found inside the dissolved oxygen sampler, with the water to be tested. Reassemble the sampler according to directions in the kit and lower the sampler into the water. At the desired depth, pull on the string to pull out the rubber stopper, thus filling the sampler and the sample bottle with water. Record the depth and location where you obtained the sample of water.
2. Raise the sampler, being careful to keep it upright. Remove the lid to the sampler and take out the completely filled sample bottle. Immediately measure and record the temperature of the water that remains inside the sampler.
3. In this step and step 4, chemicals from the kit are added to the water in the sample bottle and mixed thoroughly. The chemicals react with any oxygen dissolved in the water to produce iodine which dissolves in the water. You cannot see dissolved oxygen for it is colorless, but you can see dissolved iodine, for it is yellow or amber in color.

Add the contents of powder pillows I and II, found in the kit, to the water in the sample bottle. Tilt the sample bottle slightly and insert the glass stopper quickly. CAUTION: Be careful not to trap air bubbles in the water. Shake the bottle vigorously until the chemicals are thoroughly mixed with the water. Observe and record any changes that take place.

4. Add the contents of powder pillow III to the mixture in the sample bottle. Again replace the stopper without trapping any air and shake the bottle vigorously. Observe and record any changes that take place.

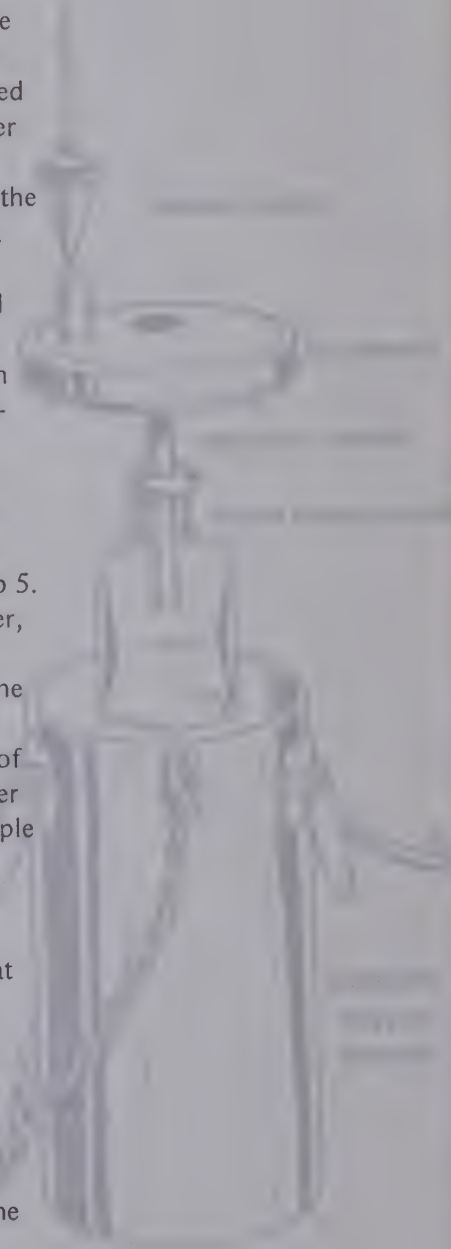
The concentration of dissolved iodine is proportional to the concentration of oxygen originally present in the water. Therefore, if no yellow color appeared at this point, there was no dissolved oxygen present in your sample and you do not need to continue the test. If, however, a yellow color did appear, proceed to step 5.

5. To determine the concentration of dissolved oxygen originally present in the water, phenylarsene oxide (PAO) solution, found in the kit, is used. PAO reacts with iodine, formed in step 4, to produce a colorless solution. The concentration of the PAO solution is such that for each drop added to a specified volume of water to produce a colorless solution, 1 ppm of oxygen was present in the original sample of water. For example, if 8 drops of PAO are added to the specific quantity of water to produce a colorless solution, 8 ppm of oxygen were present in the original sample of water.

Transfer water from your round sample bottle into the square mixing bottle as directed in the kit. While swirling the mixing bottle, add PAO drop by drop to the bottle until the water loses its yellow color; keep count of the drops. What was the concentration of oxygen in your sample of water? Record your answer.

Questions

1. How does the concentration of dissolved oxygen in your sample compare with the maximum concentration given in Table 3.
2. If the concentration of dissolved oxygen in your sample is less than the maximum concentration, what is a reasonable explanation for your experimental result? If the concentration of dissolved oxygen in your sample is more than the maximum concentration, what is a reasonable explanation for your experimental results?
3. How does a body of water regain its dissolved oxygen once it has been used up?



Suggestions for Further Study

1. Do you think the amount of dissolved oxygen is the same throughout a lake? Is it the same everywhere on the surface of a lake? Is it the same at any point along a vertical line through a lake? Check your answers experimentally.
2. Fish differ greatly in their need for dissolved oxygen. For example, at 20°C, rainbow trout require at least 2.64 ppm of dissolved oxygen, carp require 3.74 ppm of dissolved oxygen, but perch can survive on 1.25 ppm of dissolved oxygen. Therefore, the concentration of dissolved oxygen in a body of water plays an important role in determining which species of fish can live in the water. Read and report on the concentration of oxygen required by each of several other species of fish.
3. What effect will discharging warm water into a body of water have on the concentration of dissolved oxygen in the body of water? What are some alternative procedures to discharging warm water from power plants into a body of water?
4. Try to verify the information in Table 3 by determining the concentration of dissolved oxygen in several samples of distilled water at different temperatures.

ENVIRONMENTAL EXPERIMENT 2-4 • DETERMINATION OF NITRATE ION

Purpose

To determine the concentration of nitrate ion in water.

Materials

3 clean sample bottles	1 cc standard solution containing 10 ppm of nitrogen as nitrate ions
3 samples of water	1 10-cc graduated cylinder
1 glass-marking pencil	6 cc brucine sulfate solution
6 test tubes, Pyrex	30 cc 10 M sulfuric acid
1 test tube rack	glass stirring rods
standard solution containing 100 ppm of nitrogen as nitrate ions (to be used in preparing standard solutions of lower concentrations)	2 beakers, Pyrex, 600 cc
1 cc standard solution containing 1 ppm of nitrogen as nitrate ions	1 ring stand
1 cc standard solution containing 5 ppm of nitrogen as nitrate ions	1 ring
	1 wire gauze
	1 Bunsen burner
	matches
	crushed ice

Discussion

Where do nitrate ions (NO_3^-) in water come from? Nitrogen compounds from fertilizers added to soil are converted into nitrate ions by microorganisms in the soil. In addition, the decay of plant and animal matter in soil also produces nitrate ions. Surface and ground waters carry nitrate ions into bodies of water. Moreover, nitrate ions are among the products of decay of plant and animal matter in the water.

Nitrate ions are one of the required nutrients for growth of plants and are, therefore, essential for water and soil to be of good quality. Sometimes, however, the concentration of nitrate ion or of other nutrients in a body of water such as a lake becomes too high.

When other nutrients are in adequate supply the excess nutrient stimulates an abundant growth of algae, called an algal bloom. Just as an algal bloom appears rapidly, the algae die rapidly. The decomposition of an excess of dead algae adds more nutrients to the lake. It also uses up the dissolved oxygen in the water, causing fish kills. Any further decay takes place without oxygen, creating unpleasant and toxic gases.

The process by which a lake is overfertilized is called *eutrophication*. Eutrophication, which occurs naturally as a body of water ages, is accelerated by man's addition of compounds to land and water that yield nitrate ions, phosphate ions, and other plant nutrients. The algal blooms resulting from eutrophication are the main reason excess nitrate ions have been identified as a major pollution problem in some bodies of water.

There is another reason the concentration of nitrate ion in water is important. Excess nitrate ions in drinking water can have a harmful effect on animals, especially human beings. Infants in particular are sensitive to high concentration of nitrate ion since their digestive systems readily convert nitrate ions to nitrite ions (NO_2^-). Nitrite ions combine strongly with hemoglobin and prevent proper transport of oxygen by the blood. Infant nitrite-hemoglobin poisonings, called infantile methemoglobinemia, have been reported in many parts of the world. Usually the infants have been drinking well water that has a high concentration of nitrite ions.

The possible combination of nitrite ions with some of the amino acids, present in all protein matter, to form a class of compounds known as nitrosoamines is still another reason the concentration of nitrate ion in water is important. Nitrosoamines are known to be capable of causing cancer in some organisms.

In drinking water, 45 ppm of nitrogen as nitrate ions is the standard set by the U.S. Public Health Service. The Food and Drug Administration has set a standard of 500 ppm of nitrogen as nitrate ions and 200 ppm of nitrogen as nitrite ions in our food supply.

Procedure

1. Select three sources of water to sample. Here is a list of suggested sources: a well, a creek, a river, rain water, bottled water, sterilized bottled water, agricultural wastewater, and water for irrigation.

Obtain a sample of water from each of your chosen sources. You will need only 1 cc of water to measure the concentration of nitrate ion. Label each sample bottle and record the source of the water in each of the sample bottles.

2. Set up 6 test tubes in a test tube rack. You will need to use clean, dry test tubes. Label each test tube with a consecutive numeral, beginning with numeral 1.
3. A standard solution containing 100 ppm of nitrogen as nitrate ions was prepared and then diluted to obtain standard solutions containing 1 ppm, 5 ppm, and 10 ppm of nitrogen as nitrate ions. The concentration of nitrate ion in each of your samples of water should be within the range of concentrations of the standard solutions, that is, 1 ppm–10 ppm of nitrogen as nitrate ions. If after testing your samples, you decide that you need standard solutions of other concentrations, ask your teacher for permission to prepare the solutions you need and rerun the test.

To each of the first three test tubes, add 1 cc of a different one of the diluted standard solutions (See Figure 18). To each of the remaining test tubes, add 1 cc of a different one of the samples of water. Record which one of the liquids you added to each test tube.

Next add 1 cc of brucine sulfate solution, $(\text{C}_{23}\text{H}_{26}\text{N}_2\text{O}_4)_2 \cdot \text{H}_2\text{SO}_4 \cdot 7\text{H}_2\text{O}$, to each test tube. Finally, cautiously add 5 cc of 10 M sulfuric acid to each test tube.

Mix the chemicals in each test tube with a glass stirring rod. Use a different stirring rod for each test tube or wash and dry the stirring rod before using it in the next test tube.

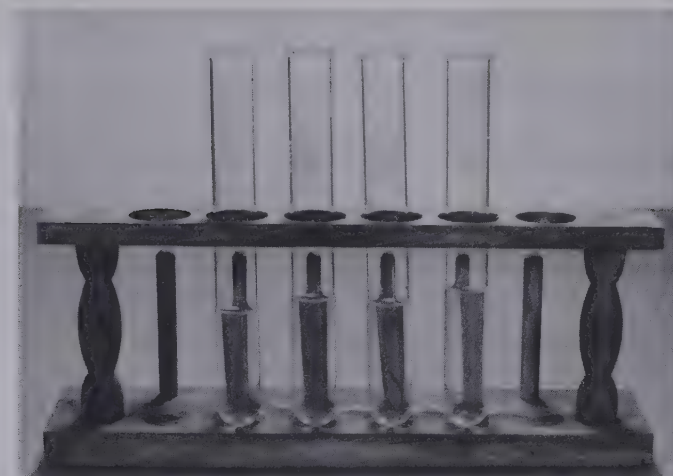


FIGURE 18 Four of the six tubes to be used in this experiment are shown. Three are for standards and the fourth represents an unknown, of which you may have quite a number. One cc of brucine sulfate goes in all tubes along with 5 cc of sulfuric acid. The standard solutions and the unknowns are the variables.

4. Place all the test tubes in a beaker of gently boiling water and heat them for 5 to 10 minutes.
5. Remove the test tubes from the water bath and put them in an ice bath to cool. When they have cooled to room temperature, compare the color of each liquid containing a sample of water whose concentration of nitrate ion is not known with the color of the liquid in each of the three test tubes containing standard solutions. Sight down through the liquid in each test tube to compare weak colors; place a sheet of white paper behind the test tubes and look through the liquid in each test tube to compare stronger colors. Record your observations.

Which standard solution most nearly matches the color of each of the liquids containing an unknown amount of nitrate ion? Estimate and record the concentration of nitrate ion in each sample of water.

Questions

1. What did you estimate the concentration of nitrate ion to be in each of your samples of water?
2. How do you account for the presence or absence of nitrate ions in each sample?

Suggestions for Further Study

1. Compounds containing nitrate and nitrite ions are commonly used as food additives. Find out why they are used and why their use is controversial.
2. Read and report on the effect artificial nitrogen fertilizers have had on the concentration of nitrate ion in ground and surface waters in agricultural areas.

ENVIRONMENTAL EXPERIMENT 2-5 • DETERMINATION OF PHOSPHATE ION

Purpose

To determine the concentration of phosphate ion in water.

Materials

nonphosphate detergent	10 cc standard solution containing 10 ppm of phosphorus as phosphate ions
distilled water	1 10-cc graduated cylinder
3 clean sample bottles	1 dropper
3 samples of water	conc. sulfuric acid
1 glass-marking pencil	red and blue litmus paper
6 test tubes, Pyrex	1 beaker, Pyrex, 600 cc
1 test tube rack	1 ring stand
standard solution containing 100 ppm of phosphorus as phosphate ions (to be used in preparing standard solutions of lower concentrations)	1 ring
10 cc standard solution containing 1 ppm of phosphorus as phosphate ions	1 wire gauze
10 cc standard solution containing 5 ppm of phosphorus as phosphate ions	1 Bunsen burner
	matches
	12 cc ammonium molybdate solution
	2 cc tin (II) chloride solution, freshly prepared
	glass stirring rods

Discussion

Where do phosphate ions (PO_4^{3-}) in water come from? Some soils are naturally high in compounds containing phosphate ions. Water running down through these soils slowly dissolves the phosphate compounds. Natural decay of plant and animal matter also contributes phosphate ions to water systems. However, agricultural runoff, municipal sewage, and industrial operations contribute much larger amounts of phosphate ions to water. Phosphate ions, like nitrate ions, are nutrients, and therefore, desirable in limited amounts. Excess phosphate ions, however, rapidly accelerate the process of eutrophication. There are no known adverse health effects from drinking water containing excess phosphate ions, and thus the Public Health Service has not established a recommended limit.

In the late 1960s, much attention was focused on phosphate ions in detergents. Since household wastes are an important component of municipal sewage, detergents having a high concentration of phosphate were thought to be a primary culprit in the visible eutrophication of bodies of water receiving either conventionally treated or untreated sewage. However, a large amount of phosphate ions (approximately 680 million tons/year overall in the United States) would still reach surface waters from the other sources mentioned above, even if all phosphate ions were removed from detergents. Nevertheless, domestic wastes in some areas provide up to 70 percent of the phosphate ions added to the water system. In such cases, the reduction of phosphate in detergents may provide a visible improvement in the quality of water.

Procedure

Clean glassware before you use it by washing it with a nonphosphate detergent and then rinsing it with distilled water. Avoid the use of tap water unless you know it is free of phosphate ions. Throughout this experiment, take special care to see that you use only clean glassware.

1. Choose three sources of water to sample. You might wish to sample agricultural wastewater, water in a creek, water in a river, or tap water. If you contact the proper official, you might be able to obtain a sample of wastewater from your local sewage treatment plant or from a local industry. However, do not sample raw wastewater before it has been treated biologically or with chlorine.

Obtain a sample of water from each of the sources you selected. You will need a 10-cc sample of water to measure the concentration of phosphate ion. Label each sample bottle and record the source of the water in each of the sample bottles.

2. Set up 6 clean, dry test tubes in a test tube rack. Label each test tube with a consecutive numeral, beginning with numeral 1.
3. A standard solution containing 100 ppm of phosphorus as phosphate ions was prepared and then diluted to obtain standard solutions containing 1 ppm, 5 ppm, and 10 ppm of phosphorus as phosphate ions. The concentration of phosphate ion in each of your samples of water should be within the range of concentrations of the standard solutions, that is, 1 ppm–10 ppm of phosphorus as phosphate ions. If after testing your samples, you decide that you need standard solutions of other concentrations, ask your teacher for permission to prepare the solutions you need and rerun the test.

To each of the test tubes labeled with numerals 1–3, add 10 cc of a different one of the diluted standard solutions. Clean the graduated cylinder each time before you use it. To each of the remaining test tubes, add 10 cc of a different one of the samples of water. Record which of the liquids you added to each test tube.

4. Because your samples of water may contain phosphorus in any of several chemical forms and this test is specific for phosphate ion, it is necessary to convert into phosphate ions other forms of phosphorus present in the water to be tested. To do this, cautiously add concentrated sulfuric acid drop by drop to each of the 10-cc samples of water (the water in the test tubes labeled with numerals 4–6) until the sample tests acidic with indicator paper. Then add 1 cc more of concentrated sulfuric acid to each of the three samples of water.

Gently heat, but do not boil, the three acidified samples in a water bath for 20 min. Note that this procedure is not needed for the standards. The phosphorus is already in phosphate ion form in them. Remove the test tubes from the water bath and return them to the test tube rack. Let the acidified samples cool to room temperature.

Check the pH of each acidified sample. It should still be acidic; but if it is not, add concentrated sulfuric acid drop by drop until the sample tests acidic.

Add distilled water, if necessary, until the surface of the acidified sample in each test tube is even with the level of liquid in the test tubes containing standard solutions.

5. To each of the test tubes in the test tube rack, add 2 cc of ammonium molybdate solution and 0.25 cc (approximately 5 drops) of tin (II) chloride solution.

Mix the chemicals in each test tube with a glass stirring rod. Use a different stirring rod for each test tube or wash and dry the stirring rod before using it in the next test tube.

Let the test tubes stand for 10 min. If the colors are slow in developing, put the test tubes in a warm water bath for 2–3 min. If the phosphate ion is present in the test tube, the color that appears is due to the formation of a complex substance called molybdenum blue.

6. Compare the color of each liquid containing a sample of water whose concentration of phosphate ion is unknown with the color of liquid in each of the test tubes containing a standard solution. Sight down through the liquid in each test tube to compare weak colors; place a sheet of white paper behind the test tube to compare stronger colors. The color will slowly begin to fade after it has reached its maximum intensity so do not delay in making your comparisons. Record your observations.

Which standard solution of phosphate ion most nearly matches the color which exists in each of the samples of water of unknown concentration of phosphate ion? Estimate and record the concentration of phosphate ion in each sample of water.

Questions

1. What did you estimate the concentration of phosphate ion to be in each of your samples of water?
2. How do you account for the presence or absence of phosphate ions in each sample?

Suggestions for Further Study

1. What is the difference in the concentration of phosphate ion if polyphosphate ions ($\text{P}_3\text{O}_{10}^{5-}$) present in the water are not converted to phosphate ions? To find out, obtain two samples of water from the same source. Measure the concentration of phosphate ion in each sample; use the test outlined in the *Procedure* but omit step 4 for one sample.
What does the difference between the concentrations of phosphate ion in the two samples tell you about the relative amount of polyphosphate ions in the water?
2. Will conventional methods of wastewater treatment remove phosphate ions from sewage waters? Does your area treat waste so as to remove nutrients?
3. Why do many detergents contain phosphate ions? Is it realistic to ban all phosphate ions in laundry and dishwashing compounds? Why or why not?

ENVIRONMENTAL EXPERIMENT 2-6 • ANALYSIS FOR LEAD ION

Purpose

To analyze for the presence of the lead ion in water.

Materials

- | | |
|--|--|
| 3 clean sample bottles | 10 cc standard solution containing 10 ppm of lead ions |
| 3 samples of water | 10 cc standard solution containing 25 ppm of lead ions |
| 1 glass-marking pencil | Hydrion paper |
| 6 test tubes, Pyrex | conc. ammonia water |
| 1 test tube rack | 1 dropper |
| standard solution containing 100 ppm of lead ions (to be used in preparing standard solutions of lower concentrations) | 30 cc dithizone solution |
| 10 cc standard solution containing 1 ppm of lead ions | 6 solid stoppers or corks to fit the test tubes |

Discussion

There are many metallic ions that have made their way into the water supply. Ions of metals such as arsenic, antimony, cadmium, chromium, copper, lead, manganese, mercury, tin, and zinc are among those that have been detected in natural waters. However, a water supply may not be routinely checked for ions of these metals even though they can be detected quite easily.

Man's activities may increase the concentration of certain metallic ions in water until they constitute health hazards. In this experiment, you will analyze water for the

presence of the lead ion. The lead ion will serve as an example of an ion of a heavy metal that is finding its way into water supplies.

Where do lead ions in water come from? Water leaches some lead compounds from natural deposits of the metal. Even so, normally natural waters seldom contain more than 0.02 ppm of lead ions. However, industrial processes, mining, and smelting operations increase the concentration of lead ions in natural waters. Furthermore, lead from lead pipes and from the solder used in copper plumbing slowly dissolves in the water flowing through the pipes, forming lead ions, especially if the water is acidic.

There is still another source of lead ions in water. It is a relatively new source, namely, fallout from automobile exhaust gases. Tetraethyl lead is added to gasoline to increase its octane rating. When leaded gasoline is burned in an automobile, a large percentage of the lead goes out the tailpipe of the automobile. As a result, the concentration of lead in the atmosphere in urban areas is increasing.

How does lead from automobile exhaust gases, as well as other airborne lead, get into the water supply? Particles of lead compounds in the atmosphere dissolve in moisture in the air forming lead ions. Rain and snow carry the lead ions to the ground. The concentration of lead ions in rainfall varies from a maximum of 300 ppm when the rain is finely dispersed to an average of 36 ppm in urban areas. Snow and melt water may contain one hundred or more times the natural concentration of lead ions. Recent studies have shown that lead ions have found their way into a remote canyon high in the Sierra Nevada in Yosemite National Park. The source of the lead ions is thought to be automobile exhaust gases carried from the San Francisco area by air currents.

Most of the lead ions carried to the ground in rain and snow find their way into the soil and then into the ground water. Some are carried by surface waters to the ocean. The concentration of lead ions in the ocean is increasing, particularly around continents having large urban areas located on or near their shores.

Lead ions are a health hazard and they may accumulate in the body. Moderate amounts cause symptoms such as anemia, headaches, and mental retardation. In very large amounts, the resulting brain damage causes convulsions, blindness, and even death. Symptoms are sometimes evident when a cold or other infection helps to release stored lead ions into the bloodstream. In recognition of the health hazards of lead ions, a limit of 0.05 ppm of lead ions in drinking water has been established.

Note: The following test should be considered as qualitative. Your sample may contain other metal ions which will react with the dithizone to give a positive test. The action of such metal ions can be partially controlled by making the solution alkaline. However, to make the test quantitative, other chemicals would have to be added to bind up the interfering metal ions. Organic matter can also interfere with this test if present, and ideally, lead-free water and glassware should be used throughout the experiment. Despite these factors, you should get a relative idea of dissolved lead present in your samples by means of this test.

Procedure

1. Select three sources of water to sample. Among the sources you might select are these: tap water, especially if lead pipes are used; water draining from a mine; water in a stream or gutter near a busy highway; rain water; and melted snow.
Obtain a sample of water from each of the sources you selected. You will need 10 cc of water to analyze for the presence of the lead ion. Label each sample bottle and record the source of the water in each sample bottle.
2. Set up 6 clean, dry test tubes in a test tube rack. Label each test tube with a consecutive numeral, beginning with numeral 1.
3. A standard solution containing 100 ppm of lead ions was prepared and then diluted to obtain standard solutions containing 1 ppm, 10 ppm, and 25 ppm of

lead ions. The concentration of lead ions in each of your samples of water should be within the range of concentrations of the standard solutions, that is, 1 ppm–25 ppm of lead ions.

To each of the test tubes labeled with numerals 1–3, add 10 cc of a different one of the diluted standard solutions. To each of the remaining test tubes, add 10 cc of a different one of the samples of water. Record which one of the liquids you added to each test tube.

Check the pH of the liquid in each test tube, using Hydrion paper. The pH should be between 8 and 9. If it is not, adjust the pH by adding concentrated ammonia water, $\text{NH}_3(\text{aq})$, drop by drop.

4. To each test tube add 5 cc of dithizone solution. CAUTION: Use dithizone solution in a well-ventilated area. Stopper and shake the test tube for 1 min. To prevent pressure from building up in the test tube, release the stopper from time to time to vent the contents. Return the test tube to the test tube rack and allow the mixture to separate into layers.

Observe the color of the lower layer in each of the test tubes. Is there a lower layer that is green or yellow in color? If a lower layer is green, the color of the dithizone solution, then the pH of the sample was too alkaline. The pH of a new sample must be adjusted between 8 and 9 before repeating the test. If a lower layer appears yellow, there may be too high a concentration of oxidizing material in your sample. Ask your teacher for an appropriate way to treat a new sample before repeating the test.

Look at the color of the lower layer in each of the test tubes containing a standard solution. Note that the color in the lower layers shades from pink to red as the concentration of lead in the standard solution increases.

Look at the color of the lower layer in each of the test tubes containing a sample of water. If a lower layer is colorless, lead is absent from the sample of water or is present in too low a concentration to be detected by this test (below 0.05 ppm). If a lower layer is pink or red in color, lead is present in the sample of water. Record your observations.

Questions

1. Do any of your samples contain lead ions? If so, what do you think are the sources of the lead ions?
2. Could you test for metallic lead by using this procedure? Explain.

Suggestions for Further Study

1. Sometimes there is a lead compound in the glaze on a ceramic container. If the lead compound leaches out of the glaze into food stored in the container, the lead ions in the compound may constitute a health hazard. For example, if there is a lead compound in the glaze of a ceramic container used to store orange juice, the glaze may slowly give up the lead compound to the orange juice, an acidic compound.

One way to check for the presence of a lead compound in the glaze of a ceramic container is to pour a solution of 5 percent acetic acid into the container, cover the container, and allow the acid to stand in the container overnight at room temperature. The next day remove the acid from the container and measure out a 10-cc sample. Neutralize the sample with concentrated ammonia water and adjust the pH to read between 8 and 9. Test the alkaline solution for lead ions using step 4 of the *Procedure*. If there are lead ions in the sample, there is a lead compound in the glaze. On the other hand, if there are no lead ions in the sample, it can be assumed that the glaze is free of lead compounds.

2. Lead poisoning is treated with chelating agents. These agents bind lead ions into compounds that are soluble. The chelated lead is then excreted through the kidney and liver. Read and report on the use of chelating agents to treat lead poisoning.
3. Lead mined from specific sources consists of a characteristic distribution of isotopes. This allows researchers to determine whether lead ions in the environment are from gasoline additives, which are only made by a few suppliers, or from other sources such as burning coal. Read and report on this chemical signature used to investigate environmental contamination by lead ions.

ENVIRONMENTAL EXPERIMENT 2-7 • DETECTION OF BACTERIAL CONTAMINATION

Purpose

To test for the presence of bacterial contamination in water.

Materials

- | | |
|---|---|
| 3 clean sample bottles | 1 wire loop in a glass-rod holder |
| 3 samples of water | 1 Bunsen burner |
| 1 glass-marking pencil | matches |
| 3 sterilized test-tube assemblies containing lactose broth made from beef extract, peptone, and lactose | 1 sterilized test-tube assembly, containing brilliant green lactose bile broth made from peptone, lactose, dehydrated oxgall, and brilliant green, for each positive presumptive test |
| 1 10-cc graduated cylinder | |
| 1 constant temperature oven or incubator | |

Discussion

The presence of disease-causing (pathogenic) bacteria in water can make it unsafe for drinking. Even though there may be many pathogenic bacteria in drinking water, they are difficult to detect by simple procedures. Therefore, coliform bacteria, which can be detected more readily, are used as an indicator of the presence of pathogenic bacteria. Coliform bacteria themselves do not commonly cause disease. Rather, they live harmlessly in the intestines of human beings and animals. Coliform bacteria, therefore, are normally associated with human and animal wastes that may also contain pathogenic bacteria. By testing for coliform bacteria, one can obtain an indication of the overall bacterial quality of the water.

One complication in carrying out tests for coliform bacteria is the uneven distribution of bacteria in a sample of water, even though the sample has been adequately mixed. For example, suppose a sample of water contains one coliform organism per cubic centimeter, the maximum set by the Public Health Service. It has been shown that about 35–40 percent of one-cubic-centimeter portions taken from that sample will test negatively for coliform bacteria just because of the irregular distribution of bacteria in the sample. For this reason, five or more portions of each sample should be tested for coliform bacteria to minimize misleading results.

Procedure

Ideally, you should test for coliform bacteria within one hour after collecting the sample of water. If you cannot begin the test within 24 hours, the sample is no longer useful. Because of this, it is recommended that you use a coliform testing kit so you can test for coliform bacteria in the field.

1. Choose three sources of water to sample. You might wish to sample tap water, well water, agricultural wastewater, water in a river, water in a bird bath, or rain water in a puddle. If you contact the proper official, you might be able to obtain a sample of wastewater from your local sewage treatment plant or from a local industry. However, do not sample raw wastewater before it has been treated biologically or with chlorine.

Obtain a sample of water from each of the sources you selected. Check your coliform testing kit to see how large a sample you will need to test for coliform bacteria. Label each sample bottle and record the source of water in each of the sample bottles.

2. Coliform bacteria can live and grow in a medium called lactose broth. Lactose broth, which is used as a food by coliform bacteria, contains beef extract, peptone, and lactose, the sugar present in milk. As coliform bacteria feed on lactose broth, a gas is produced. The appearance of the gas indicates the presence of the coliform bacteria.

Transfer a measured volume of water from each sample to a different one of the test-tube assemblies. Check the instructions in the coliform testing kit for the volume of water to transfer. Place each of the test-tube assemblies in a constant temperature oven or incubator set at 35°C.



FIGURE 19 The bottle on the right is the test assembly as it looks initially. The bottle on the left represents a positive test.

After 24 hr, examine each test-tube assembly. Is there any gas in the test tube? Are bubbles of gas rising in the test tube? Is the medium cloudy? Record your observations. Gas in the test tube, bubbles of gas rising in the test tube, or a cloudy medium constitute a positive presumptive test—that is, you can presume there is coliform bacteria in the sample of water. Save each test-tube assembly in which there is a positive presumptive test. You will use the medium in step 3 in trying to confirm the presence of coliform bacteria.

If there is no gas in a test tube after 24 hr, continue to incubate the test-tube assembly at 35°C for another 24 hr. At the end of that time, check each of the test-tube assemblies for evidence of coliform bacteria. Record your observations. Save each test-tube assembly with a positive presumptive test. If, however, you observe no gas in a test tube, terminate the test. You can assume there are no coliform bacteria in the sample of water.



FIGURE 20 Sterilize loop in hottest part of flame (left). Transfer one loopful from positive presumptive test to a sterilized tube of brilliant green culture medium (center photos). Resterilize the loop (right).

3. Since there are some rather uncommon bacteria which, if present, also produce gas as they feed on lactose, a second test is performed which confirms the presence of coliform bacteria. To each of the test-tube assemblies in which there was a positive test, another medium called brilliant green lactose bile broth is added. This medium is made from peptone, lactose, dehydrated oxgall, and brilliant green, which is a dye.

Innoculate the brilliant green lactose bile broth in each test-tube assembly with a different medium yielding a positive test. Follow the sequence pictured in Figure 20 and described in the rest of this paragraph. First, sterilize the wire loop in a Bunsen burner flame. Then using the sterile wire loop, transfer 1 loopful of medium yielding a positive presumptive test to a medium of brilliant green lactose bile broth. **CAUTION: At no point should you come in contact with the medium yielding a positive presumptive test. Quickly close the test-tube assembly containing the brilliant green lactose bile broth. Resterilize the wire loop.**

Incubate the test-tube assemblies containing inoculated brilliant green lactose bile broth at 35°C for a maximum of 48 hr. Periodically, check the inverted test tube in each assembly for the presence of a gas and record your observations. If at any time during the 48-hr period or at the end of the period, you observe a gas in the test tube, you have confirmed the existence of coliform bacteria in the sample. That is, the formation of a gas in this medium is a positive confirming test for coliform bacteria. If you observe a gas during the 48-hr period, discontinue incubating the inoculated broth. If no gas forms during the 48-hr period, you have shown that there are no coliform bacteria in the sample of water.

4. Return all test-tube assemblies yielding positive tests to your teacher for proper sterilization and disposal.

Questions

1. Which of your samples of water contained coliform bacteria? Which did not? Use this experimental evidence to comment on the meaning of the results of your test.
2. Does the presence of coliform bacteria necessarily mean that the water is polluted? Does the absence of coliform bacteria mean that the water is bacterially safe?

3. What drawbacks are there to municipal water treatment plants running tests for coliform bacteria that may take 96 hours to complete?

Suggestions for Further Study

1. There are alternative ways to test for coliform bacteria. Read and report on some of these other methods. What are their advantages compared with the two-step method used in this experiment?
2. What method is used at your local water treatment plant to test for coliform bacteria?

ENVIRONMENTAL EXPERIMENT 2-8 • PURIFICATION OF WATER BY COAGULATION

Purpose

To evaluate the use of coagulants for the purification of water.

Materials

4 test tubes	1 test tube rack
25 cc 0.20M slaked lime (calcium hydroxide) solution	1 glass-marking pencil
7 cc 0.20M alum (hydrated potassium aluminum sulfate) solution	1 10-cc graduated cylinder
	25 cc muddy water
	1 glass stirring rod

Discussion

What is a coagulant? To answer the question, examine a reaction in which a coagulant is used. For example, if you add a solution of alum (chemically, potassium aluminum sulfate, $KAl(SO_4)_2 \cdot 12H_2O$) to a solution of slaked lime (chemically, calcium hydroxide, $Ca(OH)_2$), you will observe the formation of a coagulated product, a light fluffy solid. This form of solid is called a floc, or a flocculent precipitate. Chemically, the floc formed in the example is composed of aluminum hydroxide with water of hydration, $Al(OH)_3 \cdot xH_2O$. The alum is the chemical agent responsible for the formation of the coagulated product, or floc, and therefore, is the coagulant. The process by which the floc formed is called coagulation.

Why is coagulation an important process for improving the quality of drinking water? Most natural water supplies contain suspended particles that cause the water to have an unpleasant taste and to appear cloudy. Most of these particles would settle out if given enough time. However, the settling process can be speeded up by the addition of a coagulant to the water. The process of coagulation takes place, forming a floc. The floc sinks. The suspended particles are caught in the floc and sink with it.

Procedure

1. Demonstrate for yourself the process of coagulation. Pour 12 cc of slaked lime solution into a test tube. Slowly add 3 cc of alum solution and observe the formation of the floc.

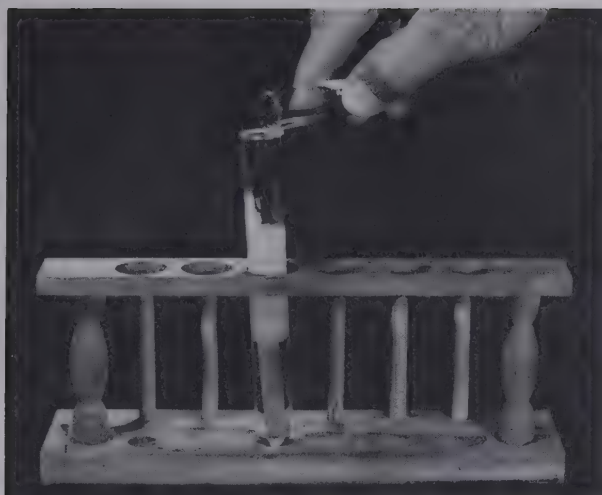
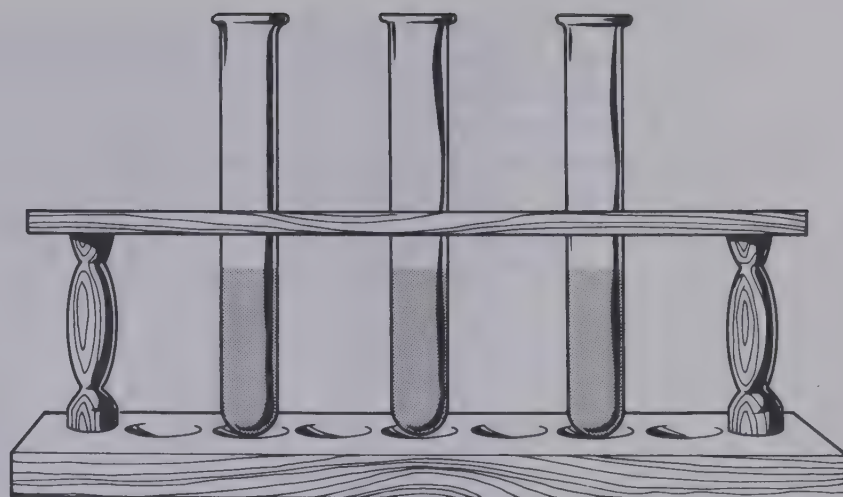
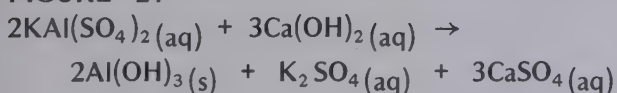


FIGURE 21



Test tube 1 Test tube 2 Test tube 3

FIGURE 22

2. Just how effective is the process of coagulation for purifying water compared with the process of simple settling? To answer this question, use the following procedure which is pictured in Figure 22.

Set up 3 test tubes in a test tube rack. Label each test tube with a consecutive numeral, beginning with numeral 1. Add 7.5 cc of muddy water to each test tube. To the first test tube add 12.5 cc of tap water and to the third test tube add 10 cc of tap water. To the second test tube add 10 cc of slaked lime solution and to the third test tube, 2 cc of slaked lime solution. To the second test tube add 2.5 cc of alum solution, and add 0.5 cc of alum solution to the third test tube. Stir the chemicals in each test tube. Wash and dry the stirring rod before using it in the next test tube. Immediately make and record your first observations. Observe and record the changes in each test tube after 5 min, after 10 min, and then at 10 min intervals until you observe no further change.

Questions

1. On the basis of your evidence, just how effective is the process of coagulation compared with the process of simple settling?
2. Write a balanced ionic equation for the formation of the floc. (You may ignore the water of hydration.)
3. What kinds of water pollutants would not be removed by the process of coagulation?

Suggestions for Further Study

1. There are many other types of coagulants in use today in addition to alum. These coagulants include iron (III) salts such as iron (III) sulfate and new organic polyelectrolytes. Devise experiments to compare their effectiveness with that of alum.
2. Repeat step 1 of the *Procedure*. Then slowly add 7 cc more of the slaked lime solution to the test tube, stirring as you add the additional slaked lime solution.

Do you see a change taking place within the test tube? If you do, try to discover the reason for the change. You may have to refer to other textbooks to find the reason.

3. Visit, telephone, or write to an official of your local water treatment facility to determine what coagulants, if any, are used in the treatment of your drinking water. Investigate the cost of the various coagulants. If coagulants are used, does the economic factor play a part in the choice of coagulants at your local water treatment plant?

ENVIRONMENTAL EXPERIMENT 2-9 • PURIFICATION OF WATER BY FILTRATION

Purpose

To evaluate the process of filtration for the purification of water.

Materials

- | | |
|---|---|
| 3 filters (suggested materials for assembling filters: funnels, filter paper of different grades, ring stands, rings, clay triangles, burette clamps, piece of cloth, glass wool, open-top separatory funnel, different grades of washed sand and gravel, Buchner funnel, filter flask, one-hole stopper to fit filter flask, thick-walled rubber tubing, and faucet aspirator) | 6 test tubes, Pyrex |
| | 1 test tube rack |
| | 30 cc muddy water |
| | 1 10-cc graduated cylinder |
| | 12 cc 0.02 M slaked lime (calcium hydroxide) solution |
| | 3 cc 0.20 M alum (hydrated potassium aluminum sulfate) solution |

Discussion

In the last experiment, you saw that the process of coagulation can be used for the removal of suspended particles from water. Once the floc has formed, however, how does one separate the water from the floc and its trapped particles? One way is to allow the floc to settle and then skim the water off the top; another way is to filter the mixture. Often these two processes are used together.

The technique of filtration has wide application in chemistry. In the laboratory you frequently use a filter known as filter paper to separate a solid from a liquid. Filters made of fabric are used to trap particulate matter in stack gases. Filters also play a role in a water treatment plant. To keep objects, such as fish and floating wood, out of a water treatment plant, a filter made of large-mesh screen is used. A filter made of different-sized rocks is often used as a final filtration step to "polish" the purified water. Moreover, in the treatment of wastewater, a filter made of a layer of rock on top of a layer of sand is used. The rock is inoculated with bacteria that digest the organic wastes.

In this experiment, you will evaluate the relative effectiveness of different kinds of filters for separating water from floc laden with particles that were suspended in the water.

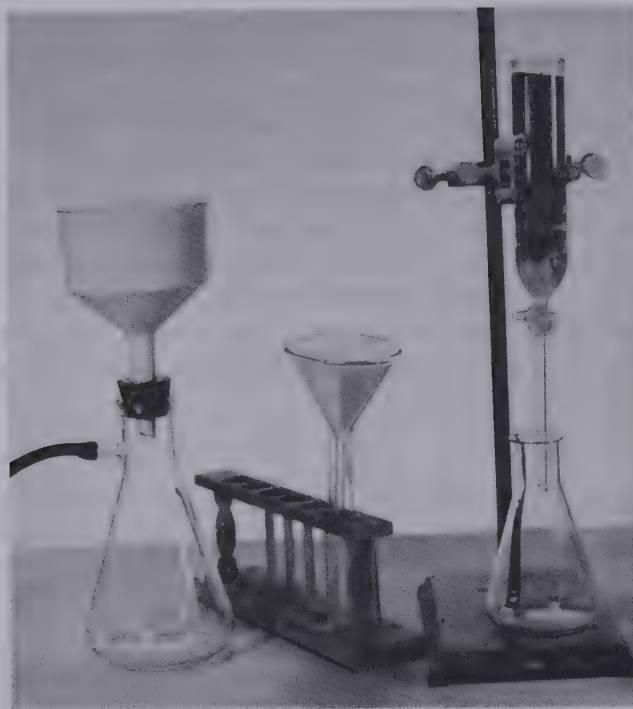


FIGURE 23 Three filtering devices. Number 1 utilizes a vacuum to draw liquid through fine pore filter paper. Number 2 uses a coarser filter paper. Number 3 uses sand and gravel particles of varying size.

Procedure

1. Select three different kinds of filters to evaluate. You might wish to evaluate filters of different grades of filter paper, a filter of glass wool, a filter of cloth, or a filter of sand and gravel. If available, you could evaluate a Buchner funnel with either a paper or cloth filter. Some of the types of filters you might evaluate are illustrated in Figure 23.

Set up the three filters you selected. If possible, include a test tube to receive the filtrate. Make a record of each setup. Use a labeled diagram, if you wish.

2. Set up 3 test tubes in a test tube rack. Into each test tube, pour 10 cc of muddy water, 4 cc of slaked lime solution, and 1 cc of alum solution. Stir the chemicals in each test tube.

Allow 5–10 min for the floc to form and settle. Slowly pour the contents of each test tube through a different one of the filters. Record your observations.

3. If you used a Buchner funnel, transfer the filtrate from the filter flask into a test tube. Stand each of the test tubes containing a filtrate in a test tube rack. Compare the filtrates on the basis of volume and appearance. Record your observations.

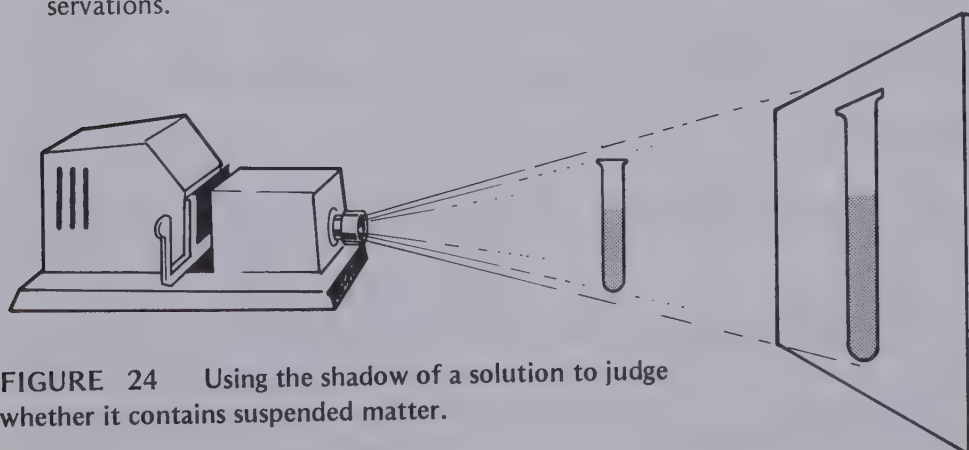


FIGURE 24 Using the shadow of a solution to judge whether it contains suspended matter.

To help you determine if the filtrate in any of the test tubes still contains a small amount of suspended matter, shine a light, such as the light from a projector, through the filtrate in each test tube. Can you see the path of the light as it passes through the filtrate? If there are suspended particles left in the filtrate, they will scatter the light, making its path visible.

An alternative method for determining whether a filtrate still contains suspended matter is this: Pour 15 cc of distilled water into a clean, dry test tube. Place a screen on the opposite side of the filtrate from the source of light. (See Figure 24) Estimate the relative darkness of the shadow, if any cast on the screen by comparing it with the shadow, if any, cast by distilled water. Should you see a shadow, it will be caused by suspended particles in the water.

Questions

1. Evaluate each filter tested for its effectiveness in removing floc laden with particles that had been suspended in the water. Which of your filters was most effective in removing the floc?
2. Before you could reuse the filters you have evaluated, you would need to clean them. Which one of the filters do you think would be easiest to clean? How would ease of cleaning enter into the choice of a filter for use on a large scale?
3. What types of materials cannot be removed from water by filtration?

Suggestions for Further Study

1. Visit, telephone, or write to an official at your local water treatment facility to determine what types of filters, if any, are used in the treatment of your drinking water. Ask the same question of an official of your local municipal waste treatment plant or an industrial waste treatment plant.
2. Activated carbon makes a highly useful filter. Not only does it have the physical ability to trap particles, but it also has the capacity to adsorb many organic chemicals such as pesticides, solvents, and harmful biological matter. It also adsorbs substances that cause water to have a taste and an odor. Does activated charcoal remove phosphate ions and nitrate ions from water? Plan an experiment to find the answer to this question. You may wish to refer to Environmental Experiment 2-4 and 2-5 for help in detecting nitrate ions and phosphate ions, respectively, in water.

ENVIRONMENTAL EXPERIMENT 2-10 • PURIFICATION OF WATER BY ION EXCHANGE

Purpose

To evaluate the process of ion exchange for the purification of water.

Materials

- | | |
|------------------------------|--|
| 1 open-top separatory funnel | 100 cc calcium chloride solution |
| glass wool | containing 0.5 g CaCl_2 per liter |
| 1 ring stand | 1 beaker, 250 cc |
| 1 burette clamp | 1 glass stirring rod |
| 3 Erlenmeyer flasks, 250 cc | 2 solid stoppers or corks to fit an Erlenmeyer flask |
| 5 g zeolite | 20 cc standard soap solution |
| distilled water | 1 dropper |
| 1 10-cc graduated cylinder | |

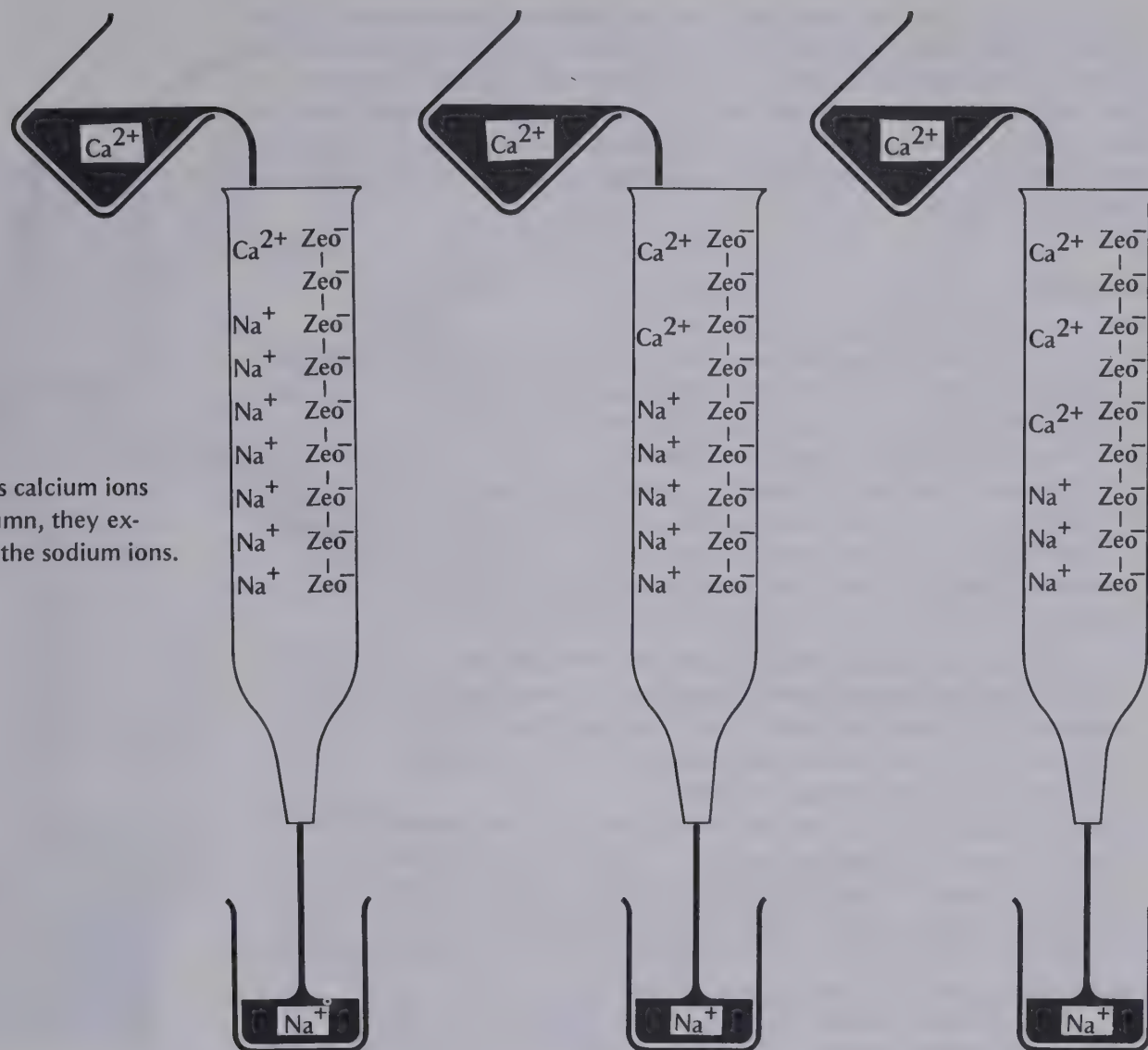


FIGURE 25 As calcium ions flow down the column, they exchange places with the sodium ions.

Discussion

Ion exchange is a chemical method used to remove certain ions from ionic compounds dissolved in the water. When water with dissolved ionic compounds flows through an adsorbent material, ions from the adsorbent material are exchanged for ions in the water. One of the earliest applications of the ion-exchange method was the removal of calcium ions from hard water by exchanging them for sodium ions. To accomplish this exchange, hard water is passed through a column filled with a natural silicate mineral (or its synthetic equivalent) known as zeolite. The approximate formula for zeolite is $\text{NaH}_6\text{AlSiO}_7$. (Na^+Zeo^- is used as an abbreviation for the formula in Figure 25. Zeolite forms a network in the column. The network is made up of large, essentially fixed, negatively charged zeolite ions and small, more mobile, positively charged sodium ions. When hard water containing calcium ions flows through the zeolite, the calcium ions, which are positively charged, are strongly attracted to the negatively charged zeolite ions, displacing sodium ions and resulting in an exchange of ions—two sodium ions for each calcium ion. (See Figure 25)

There are many types of adsorbent materials other than zeolite used in ion-exchange units. One of the newer materials consists of a synthetic hydrocarbon network containing either covalently bonded $-\text{NH}_3^+$ or $-\text{SO}_3^-$ groups. An adsorbent material can be selected that will exchange a more desirable kind of ion for the kind of ion to be removed. Further control of an ion exchange process can be achieved

by adjusting the conditions under which the ions are exchanged, including pH, type of solvent, temperature, and pressure. By combining types of resins, both positively charged ions and negatively charged ions can be removed, producing essentially ion-free or demineralized water. Other uses of ion-exchange columns include the treatment of acid wastes from mines, treatment of brackish water, separation of rare-earth metals, removal of salts from pharmaceutical preparations, and the concentration of uranium from its ores.

Procedure

1. Set up a small ion-exchange column using zeolite (Figure 26). To do this, first put a plug of glass wool in the bottom of an open-top separatory funnel. Then attach the separatory funnel to a ring stand with a burette clamp and set an Erlenmeyer flask under the separatory funnel.

Now put a few grams of zeolite in a beaker. (Use enough zeolite to make a column 10 cm high in the separatory funnel.) Add 10–20 cc of distilled water to the zeolite and stir the mixture to form a slurry. Transfer the slurry in small portions to the separatory funnel. Tap the sides of the funnel to allow the zeolite to settle evenly.

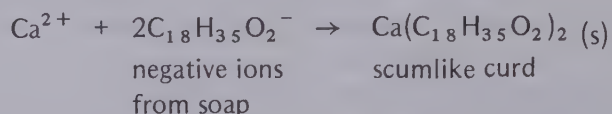
When the zeolite has settled, open the stopcock; allow water to drain off until its upper level is just above the top of the column of zeolite, and then close the stopcock. Discard the liquid in the Erlenmeyer flask and return the flask to its original position beneath the separatory funnel. The column can now sit undisturbed until you are ready to use it.

2. Add a small portion of the calcium chloride solution to the separatory funnel, using the procedure shown in Figure 27. (Adding the calcium chloride solution in this way leaves the top of the zeolite undisturbed.) Partially open the stopcock and allow the liquid to drain slowly through the zeolite. Continue to add calcium chloride solution in small portions, keeping the level of the liquid in the separatory funnel above the zeolite at all times.

When you have collected a volume of liquid approximately equal to that of the packed column, replace the Erlenmeyer flask with a clean, dry one and discard the liquid in the first flask. Collect at least 40 cc of the liquid (ion-exchange solution) in the second flask. Then close the stopcock and discontinue adding calcium chloride solution to the separatory funnel.

3. How would you know if the ion-exchange column is really working? You could test the solution for calcium ions both before and after it flowed through the column. If the zeolite is removing calcium ions from the solution, there would be fewer calcium ions in the liquid that drains from the separatory funnel than there were in the solution before it ran through the zeolite.

One convenient way of testing for calcium ions is to make use of the fact that calcium ions in hard water react with negative ions from soap molecules to form an insoluble scumlike curd. No lasting suds can be produced until all of the calcium ions are tied up with the negative ions according to this reaction:



Any soap added after the calcium ions have been tied up will produce suds (a foam). Therefore, by measuring the amount of soap needed to produce a stable foam, you can obtain an indication of the relative amount of calcium ions in the water.

To carry out this test, pour 30 cc of the original calcium chloride solution and 30 cc of your ion-exchanged solution into different clean, dry Erlenmeyer

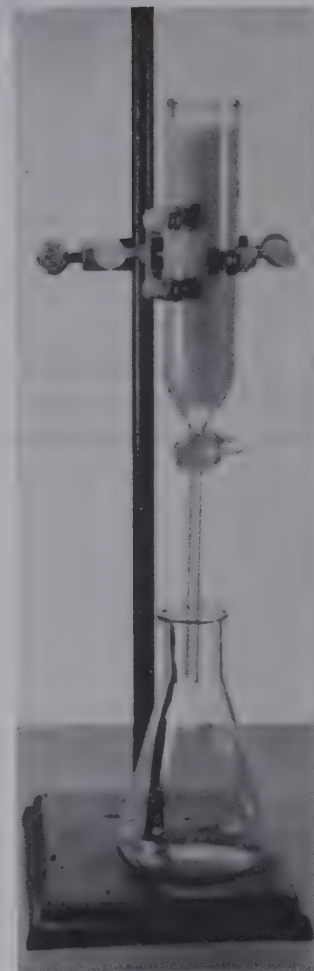


FIGURE 26 A zeolite resin column ready for use.

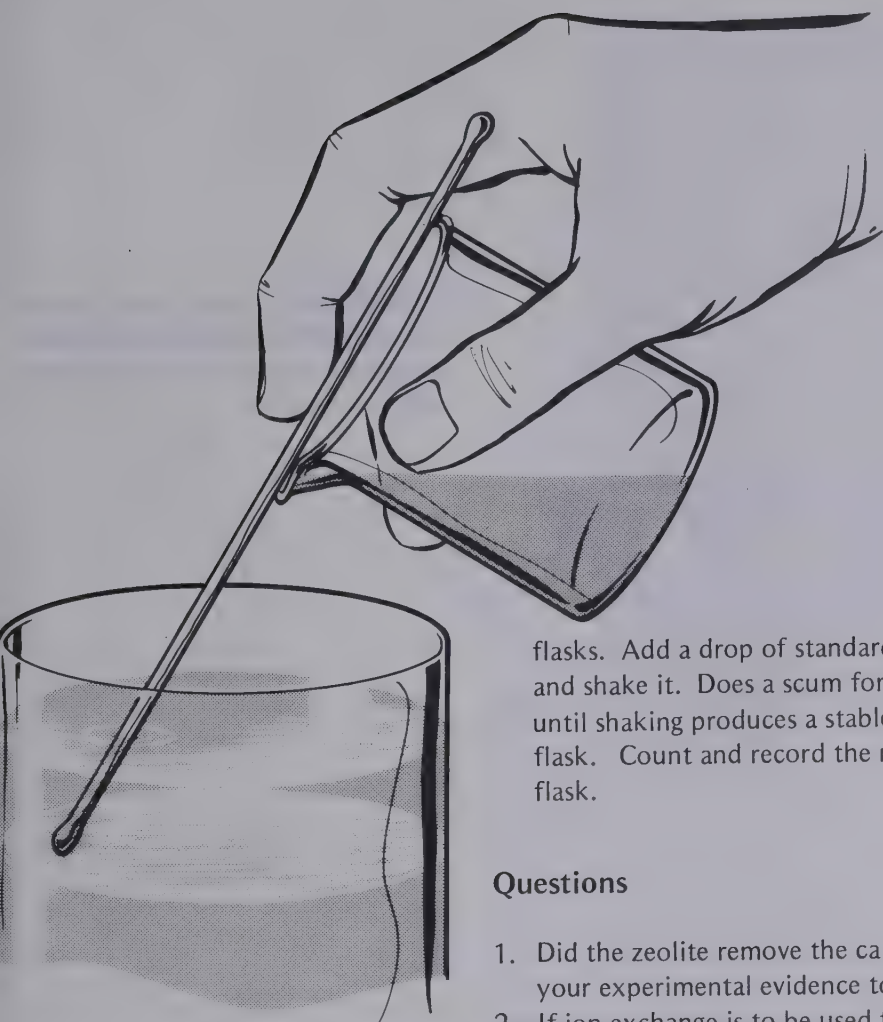


FIGURE 27 Pouring technique to use so as to avoid splashing.

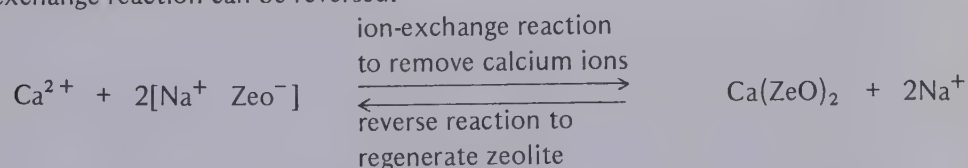
flasks. Add a drop of standard soap solution to each flask. Stopper the flask and shake it. Does a scum form? Does a foam form? Repeat the procedure until shaking produces a stable foam—a foam that lasts at least 1 min—in each flask. Count and record the number of drops of soap solution added to each flask.

Questions

1. Did the zeolite remove the calcium ions from the calcium chloride solution? Use your experimental evidence to discuss this question.
2. If ion exchange is to be used for the purification of drinking or waste waters, how must the water be pretreated?
3. What happens to the chloride ions in the original calcium chloride solution in this process of ion exchange?

Suggestions for Further Study

1. Carry out ion exchange using commercially prepared water-softening resins, such as those sold for home use. Determine the relative amount of calcium ions in the original water solution and in the ion-exchange solution.
2. When a zeolite column has been used for some time, all the sodium ions available for exchange will have been used up. To regenerate the column, the ion-exchange reaction can be reversed.



This can be done by soaking the column in concentrated sodium chloride solution. The excess of sodium ions causes the reverse reaction to occur.

Carry out a regeneration of a highly used zeolite column and rerun steps 2 and 3 of the *Procedure* to see if the effectiveness of the column has been restored.

Explain why people add rock salt to their home water softeners.

3. Invite a representative of a local water softening and conditioning company to visit your class and explain the techniques used for water treatment, or arrange to visit such a company, or do both.

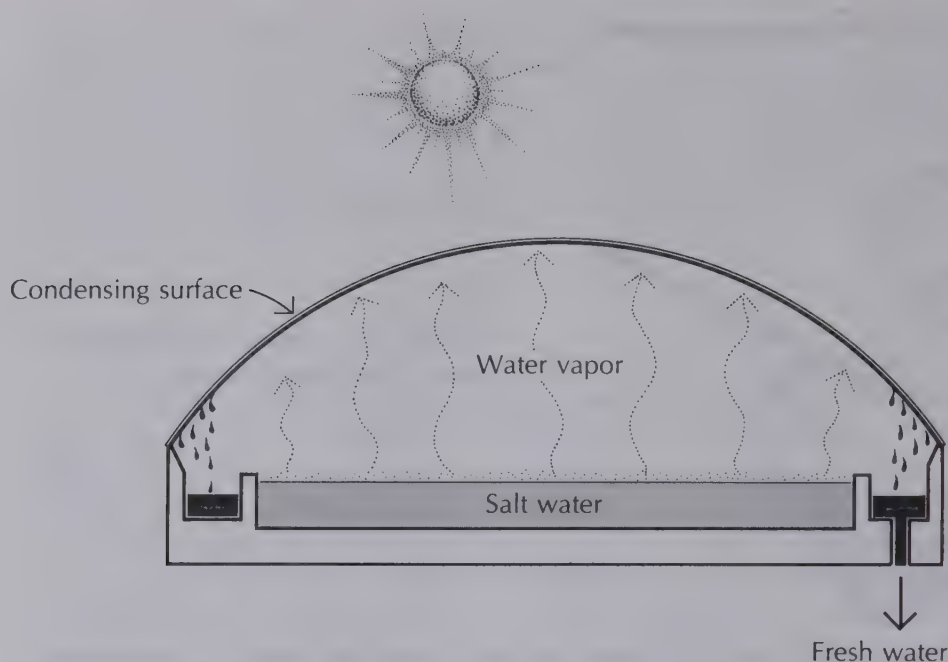


FIGURE 28 Solar humidification uses the sun's energy to obtain fresh water from salt water by evaporation and condensation.

ENVIRONMENTAL EXPERIMENT 2-11 • PURIFICATION OF WATER BY DISTILLATION

Purpose

To evaluate distillation as a method for purifying water.

Materials

- | | |
|--|---|
| 2 ring stands | 1 Erlenmeyer flask, 250 cc |
| 2 burette clamps | 1 Bunsen burner |
| 1 ring | 50 cc of dil. copper (II) sulfate solution |
| 1 wire gauze | 1 10-cc graduated cylinder |
| 1 Celsius thermometer | boiling chips such as marble chips or glass beads |
| 1 distilling flask, 250 cc | 50 cc dil. ammonia water |
| 1 one-hole stopper to fit distilling flask | phenolphthalein solution |
| 1 condenser | 1 test tube, Pyrex |
| 2 pieces of rubber tubing, 60–100 cc long | |

Discussion

Compared with water, salt solutions have higher boiling points. Since antiquity, people have separated water from salts dissolved in it by vaporizing the water and leaving the salts behind. The water vapor can be condensed and collected. This process, which is called distillation, has many different applications. One of these is simple-distillation. If enough energy is added to actually make the water boil, simple distillation occurs. This is a two-step process, involving both boiling (rapid evaporation) and condensation. Figure 29 shows the apparatus you will use to carry out simple distillation in the laboratory.

Another application of the process of distillation is solar humidification. In solar humidification, the water does not boil. Rather, energy from the sun is used to heat the water, causing it to vaporize much more rapidly than it normally would. The container holding the vaporizing water is covered by an arched plastic or glass shield (See Figure 28). The shield is cooler than the water because the material it

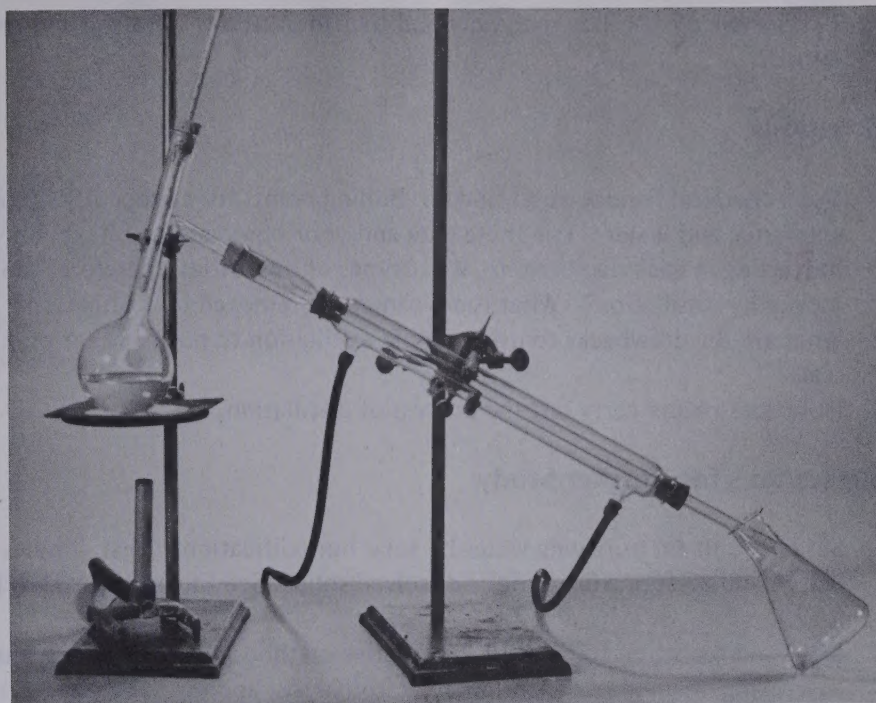


FIGURE 29 A typical laboratory set up for distilling a small amount of liquid.

is made of does not absorb heat as well as the water and its container. When the rising water vapor comes in contact with the cooler shield, it condenses. The water flows down the inside of the arched shield and is drained off.

To make it economically feasible to use the process of distillation for the purification of water on a large scale, steps must be taken to minimize the energy required to heat or boil the water. For example, in the multistage flash-distillation process for distillation of sea water, energy is minimized in several ways. The process makes use of the fact that water boils at a lower temperature if the pressure is lowered. Heated sea water passes through a series of chambers with successively lower pressure, causing very rapid, or flash, vaporization. A further saving of energy is realized by using the cool incoming sea water as a cooling agent in the condensers. The incoming sea water is itself, thereby, heated, resulting in a still further saving of energy. Furthermore, waste heat from a power plant on the same site may be used to supply the additional energy needed to heat the sea water. Plants using the multistage flash-distillation process presently account for 98 percent of the world's capacity to desalinate sea water.

Procedure

1. Set up the apparatus as shown in Figure 29. Add 30–50 cc of dilute copper (II) sulfate solution to the distilling flask. The copper (II) sulfate solution will serve as an example of a water solution of a nonvolatile solid. Observe and record the initial appearance of this solution. Also add a few marble chips or glass beads to the distilling flask to prevent uneven boiling or bumping.
2. Gently heat the distilling flask until a small volume of liquid (the distillate) accumulates in the receiving flask. Observe and record the temperature of the vapors as they leave the distilling flask. Compare the appearance of the distillate with that of the original solution. Record your observations.
3. Empty the distilling and receiving flasks and clean them. Dry the receiving flask. Reassemble the apparatus. Add 30–50 cc of dilute ammonia water to the distilling flask along with a few marble chips or glass beads. If the odor of ammonia water is not already apparent to you, cautiously observe and record its odor. Add a few drops of phenolphthalein solution to a separate 1-cc sample of dilute ammonia water. Observe and record the results.
4. Repeat step 2. Observe and record the odor, if any, of the distillate. Test a

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1-cc sample of distillate with phenolphthalein solution. Observe results.

Questions

1. Use a chemical handbook to find the boiling points for copper (II) sulfate, ammonia, and water. Use these data and your observations to explain what is happening in each distillation. What types of impurities, therefore, can be removed by distillation? What types cannot be removed by distillation?
2. What are the drawbacks to using simple distillation to purify water on a large scale?
3. How does nature carry out the process of distillation?

Suggestions for Further Study

1. Set up a unit for purifying water by solar humidification. Test samples of water before and after purification for dissolved solids (See Environmental Experiment 2-2.).
2. Read and report on some of the alternative methods of desalting sea water, such as freezing, membrane separation, electrodialysis, and ion exchange.
3. If sea water were contaminated with oil, what effect would the oil have on the flash-distillation process?
4. Could you use simple distillation to recover gold from sea water?

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This booklet will give you more information about the special properties of water, the sources and characteristics of water pollutants, and some of the major problems in and the prospects of achieving quality water. The book includes a glossary of environmental and chemical terms.


U.S. Department of the Interior, Office of Saline Water, *New Water, The A-B-Seas of Desalting*, and *Saline Water Conversion Summary Report 1972-1973*, Washington: U.S. Government Printing Office, 1970, 1968, 1973, respectively. (Document 0-381-844, 36 pages, \$0.60; Document 0-306-916, 40 pages, \$0.35; and Document 0-488-883, 70 pages, \$1.25, respectively) The first booklet gives general information about the goals of the Office of Saline Water and how they have been implemented. The second booklet gives more details of the technical processes used to produce fresh water from saline. The third is a report issued yearly discussing the programs of the Office of Saline Water. It includes details on the current status of research and development of desalination.

CONCLUSION

Now that you have studied air and water quality with the guidance of this module, it seems like a good time to step back and see what you have accomplished. You have taken an important first step in helping to bring about environmental change, namely, you have acquired knowledge about the chemical basis of air and water pollutants and of techniques for purifying air and water. Having measured the amount of various pollutants in air and water and the relative effectiveness of a variety of methods for purifying air and water, you are in a far better position than many people to understand the basis for environmental issues related to air and water quality. Now, what can you do to utilize this knowledge? Here are some suggestions for you to consider.

1. Summarize what you have learned about the quality of the air and water where you live. What issues related to air and water quality are of special concern in your area? What solutions are possible?
2. Communicate your information on a topic of environmental concern to others by writing a letter to the editor of your school or community newspaper.
3. Give an oral report to your chemistry, English, or speech class on a topic you have explored in this module. Give the same report to a group in your community.
4. Use photography or art work as a tool for communicating to your classmates and others your understanding of the importance of air quality and water quality. Pictures showing air and water pollution may help others to understand the need for and the benefits from improving air and water quality.
5. In general, people have been slow to understand the environmental effects of a life style that requires a high consumption of energy and emphasizes personal convenience through the use of disposable items. Formulate a personal code of environmental ethics which, if adopted by many people, would change the life style and result in cleaner air and purer water as well as a healthier environment for everyone.
6. Increase your effectiveness to bring about environmental change by joining an organized group. If you have unusual stamina, organize your own group to work on a local environmental problem.
7. Continue your study of environmental chemistry. This may mean going into greater depth on a topic already touched upon or branching out into related fields. For example, look into the chemistry of food additives, pesticides, plastics, or nuclear energy. All of these create problems in your external or internal environment—problems that can be better understood through chemistry.

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